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Practical Chemistry,
WITH NOTES AND QUESTIONS
ON
THEORETICAL CHEMISTRY.



WILLIAM RIPPER.

Practical Chemistry,
WITH NOTES AND QUESTIONS
ON
THEORETICAL CHEMISTRY.

*ADAPTED TO THE REVISED SYLLABUS
OF THE SCIENCE AND ART DEPARTMENT
FOR THE ELEMENTARY STAGE OF INORGANIC CHEMISTRY.*

BY
WILLIAM RIPPER,
(Member of the Cleveland Institution of Engineers),
SCIENCE MASTER IN
THE SHEFFIELD SCHOOL BOARD'S CENTRAL HIGHER SCHOOL.

IN TWO PARTS

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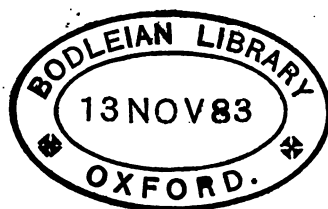
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196.



P R E F A C E .

The following pages, intended in the first place for the students in the Sheffield School Board's Central Higher School, are submitted to teachers and students of chemistry, with the hope that they may be of service in other similar classes in preparing for the Science and Art Department's Examinations.

The work is divided into two parts :—

- I. A course of Elementary Practical Chemistry.
- II. Notes and Questions on the Theoretical course of elementary Inorganic Chemistry, intended to supplement the student's manuscript notes of Lectures and to be used with or without a text book.

The experiments herein described have been carefully selected, and are such as may be conveniently performed by young students in a laboratory. Many practical hints are given on the use of the apparatus, and for the careful performance of the experiments.

The reactions of the metals required by the Department's syllabus are arranged in a tabular form, so that the various tests can be much more readily compared and remembered than by the ordinary arrangement of reactions.

A large number of the Questions have been selected from Science and Art Department and other public Examination papers; and the problems have been arranged to ensure a knowledge of principles rather than as long Arithmetical Exercises.

The necessity for a large amount of practice in working chemical problems, especially in the earlier stages of the study, is admitted by all teachers of chemistry, and it is hoped that the solutions of typical examples, and the various hints given, will enable the student, after a little thought, to answer any of the Questions without difficulty.

The *Notes and Questions* have been in the hands of my pupils for some time, and to their use is largely attributed the success which has attended the Day Classes (composed of boys from 12 to 14 years of age) held at the Sheffield Central School. At the past two examinations held by the Science and Art Department in May, 1881 and 1882, 218 papers have been worked in the Elementary Stage of Theoretical and Practical Chemistry inclusive. Not one student failed in either subject, and 60 per cent. obtained first class certificates and prizes.

The Department's "List of Experiments for the guidance of teachers" and some blank leaves for notes are inserted at the end of the book.

I am indebted to Mr. GEORGE HIRST, Assistant Science Teacher in this school, for having kindly read and corrected my manuscript and the proof sheets.

Central Schools, Sheffield.

W. R.

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PART I.

PRACTICAL CHEMISTRY.

P A R T I.

AN ELEMENTARY COURSE OF
PRACTICAL CHEMISTRY.

1. INTRODUCTORY.

(a) "Strict order, great cleanliness and neatness," are indispensable to success in Practical Chemistry.

(b) Before experimenting, find out exactly what you are going to do, and what result you are to expect, and have everything ready beforehand.

(c) If the result of your experiment is not what it should be, rest assured that the fault is with the experimenter. Some error has been committed, or something necessary to the successful result has been omitted.

The laws of Nature are unchangeable ; a stone, when thrown up, always falls to the ground ; and the facts of Chemistry have been established by long and repeated trial and experiment.

(d) The experiments in the elementary stage of Practical Chemistry are mostly confined to obtaining certain results, without reference to the *actual weights* of the quantities taking part in the reactions. *Quantitative* experiments follow, after some knowledge of *qualitative* Chemistry has been obtained.

It is most important, however, to bear in mind that in every chemical change, whether it be the combination of elements to form a compound, or the splitting up of a compound into its component elements, the changes take place according to a definite fixed proportion *by weight*.

It is especially to impress this fact upon young students that the numbers representing the weights of the substances are given with many of the equations.

It should be noted that the totals on both sides of the equations are, of course, invariably equal. Numerous exercises on these are given in Part II. Read Question 357, Part II.

2. Hints on the Construction and Use of Apparatus.

(a) *To bore a hole through a cork.*—Use a cork borer a shade less in diameter than the glass tube to be fitted; wet the cork borer; begin at the small end of the cork; push and twist the borer; notice as you proceed by looking along the axis of the cork, and at the same time turning it round, that the cork borer is going through parallel with the axis of the cork in any position. Remember, the hole has to be *cut*, not *forced* through. Grease or wet the glass tubing before pushing it into the cork; twist the tube as you push it, and do not be in a hurry. The glass tubing should be held in a duster or towel, in case it should break.

(b) *To cut glass tubing.*—Make a little notch in the glass with the edge of a file by drawing the file (not too hard) across the tube two or three times. Break it by laying it on the thumbs with the notch outwards and between the thumbs, and giving a sudden gentle pull of the fingers and push of the thumbs. The sharp edges should be rounded a little by heating with the blow-pipe flame.

(c) *To bend glass tubing.*—Hold the tubing in the flame of an ordinary batwing gas burner—this flame enables us to heat a greater length of tubing at once than with the Bunsen flame. Keep the glass towards the top of the flame, and turn it round with both hands till it softens; then remove the tube from the flame, and allow it to bend almost of itself to the required shape.

(d) *To “draw out” glass tubing to a jet.*—Hold the piece of tubing in both hands, and heat it at a point between the hands in the upper part of the Bunsen flame; turn it round until soft, then remove it from the flame and draw the two ends apart.

(e) Corks must be carefully fitted; if your cork leaks it will be better to bore another than to waste time patching it up.

(f) Glass vessels to be heated should be made of thin glass. Thick glass vessels invariably crack when heated, unless great care is taken. Always use wire gauze or a sand bath when heating glass flasks or beakers over the Bunsen flame.

(g) Thistle funnel tubes (*see* Hydrogen Apparatus), unless they have bulbs, must be placed *under* the liquid in the flask. Delivery tubes should not come far below the cork, and must, of course, be well out of the liquid.

To make a Wash Bottle.—Take a clean wide-mouthed flask holding about a pint, and fit to it a sound cork. You will now require some glass tubing about $\frac{1}{4}$ -inch outside diameter. With a

cork borer of suitable diameter bore two holes through the cork parallel to one another. Cut off a piece of tubing equal in length to twice the height of the flask. About two inches from one end draw it out to a jet. Cut off the end and arrange to have a small hole through the jet after rounding the sharp edges of the glass in the flame. Now bend the tube as shown in Fig. 1 so that when in its place and reaching nearly to the bottom of the flask the bend shall be just above the cork. For the second hole, bend another piece of tubing as shown.

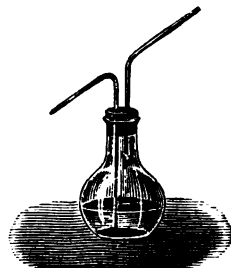


FIG. 1.

Nearly fill the bottle with distilled water. If now the student blow through the short tube, the pressure on the surface of the water will force the water up the long tube and out through the jet.

The apparatus is used for washing precipitates on filters, and as a store of warm water.

3. Preparation of pure Distilled Water from Salt Water.



FIG. 2.

Fit up an apparatus as in Fig 2. Dissolve some salt in water in a flask and pour first a little of the salt water into a test tube. Add to the solution in the test tube a few drops from the bottle labelled *silver nitrate*; a dense white curdy precipitate of *silver chloride* is formed which dissolves, when ammonia solution is added

to it. You will soon be familiar with this test, as it is always used to detect the presence of chlorides (common salt=sodium chloride). Pour the remainder of the salt water into the bulb of the retort; see that the retort is dry on the outside; use the wire gauze and apply heat. The water boils and is converted into steam, which passes over and is condensed in the flask. The flask fits loosely in the neck of the retort, and is kept cold by allowing a stream of cold water to flow over it. The boiling must not be continued to dryness in the retort, but when the liquid has reached a small bulk, pour out into an evaporating dish and set aside to cool. Crystals of sodium chloride separate out.

Take the flask containing the freshly distilled water and pour out into a *perfectly clean* test tube. Now add to this a few drops of silver nitrate; no precipitate appears, showing that the distilled water which you have obtained is perfectly free from the salt which was previously dissolved in it.

Sailors carry a distilling apparatus and in a similar way obtain perfectly pure water from sea water. Rain is distilled from the sea by the heat of the sun.

4. Illustration of the Formation of a Chemical Compound from a Mechanical Mixture.

Place some finely powdered sulphur in a small flask and drop in some bright copper turnings on the sulphur. We now have a mechanical mixture of copper and sulphur.

Place the flask on a tripod stand, and heat carefully over the Bunsen flame. The whole apparatus should stand in your pneumatic trough containing a little water in case the flask should break.

The sulphur first melts, then blackens, and at length boils. The particles of copper now become red hot, and immediately vigorous chemical action commences; the burner may be removed, the copper and sulphur are uniting together to form a new compound, and in doing so give out intense light and heat, and glow with great brilliancy.

This experiment is a type of most of the experiments to be performed with the elementary gases in the following pages.

When the flask is cold you will find, on examining the residue, *that the mass is black*, and resembles neither copper nor sulphur;

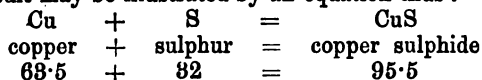


FIG. 3.

it is, in fact, a compound entirely different in its properties from either of the two substances composing it. The name given to the compound is *Copper Sulphide*.

The copper unites with the sulphur, not in irregular or varying quantities, but in a certain fixed and definite proportion by weight. The proportion is 63.5 parts by weight of copper to 32 parts by weight of sulphur. This mixture when heated gives a total of 95.5 parts by weight of copper sulphide.

This result may be illustrated by an equation thus :—



The numbers 63.5 and 32 represent the *atomic weights* of copper and sulphur and these numbers are fixed and invariable for these substances.

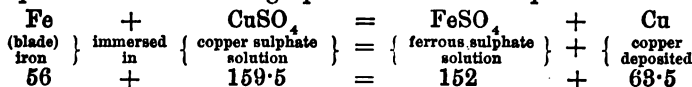
A table of the atomic weights of the elements is given on page 63, Part II.

Any excess of, say sulphur, in the above experiment, would remain in the flask uncombined, and in that case we should then have a mixture of copper sulphide and sulphur.

5. To Show the Presence of a Metal in a Solution of Blue Vitriol.

Dissolve a small crystal of blue vitriol (copper sulphate) in a test tube half full of hot water; add one or two drops of HCl and immerse a clean knife blade in the liquid. In a few minutes the blade is coated with red metallic copper, which looks bright when rubbed. If the blade remain immersed for some time, all the copper is deposited on it, and the solution loses its blue colour. Now immerse another clean knife blade in the colourless solution; no further deposit of copper appears.

Two questions will present themselves to the student (i) Has the knife blade been affected? (ii) What is now contained in the liquid? Let the following equation answer these questions :—



The equation shows us (i) that the iron of the blade and the copper of the copper sulphate solution exchange places; and for every 63.5 parts by weight of copper deposited on the blade, 56 parts by weight of iron are dissolved to take the place of the copper; and (ii) that at the end of the experiment the solution contains ferrous sulphate (Lat. ferrum, iron) instead of copper sulphate.

The number 159.5 is obtained by turning to the Table of Atomic Weights and adding together the weights of copper, of sulphur, and four times the weight of oxygen = $63.5 + 32 + (16 \times 4)$
 = $63.5 + 32 + 64$
 = 159.5

The number 152 is obtained similarly.

To prove whether iron really has been dissolved in the solution, add to it a few drops from the bottle on your bench labelled *Potassium Ferrocyanide*. If iron is present you will see a light blue precipitate which rapidly turns to dark blue; if copper is present the precipitate formed is brown.

6. HYDROGEN (H).

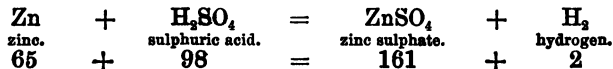
Preparation. By the action of dilute Sulphuric or Hydrochloric Acid on Zinc.

Fit up the apparatus as in Fig. 4.



FIG. 4.

Put some granulated zinc carefully in the flask (hold the flask on its side or you may knock the bottom out); pour on water till the zinc is covered, and add a little sulphuric acid by the thistle-funnel tube. Zinc sulphate is formed and hydrogen is given off, at first mixed with air.



Collect a test tube full and apply a light; it explodes sharply. The hydrogen united explosively with the oxygen of the air contained in the flask to form water. Continue to collect the gas in a test tube, until, on applying a light, it burns quietly.

All the air has now been driven out of the flask, and the hydrogen is pure. Collect three jars of the gas.

Jar 1.—Apply a lighted taper; the gas burns with a blue flame, and in doing so forms water. Notice the deposit of moisture on the jar.

Jar 2.—Pour the hydrogen *upwards* into a jar of air held with its mouth over the jar of hydrogen (Fig. 5.) Apply a lighted taper to the *upper* jar; an explosion occurs; the lighter hydrogen has risen and mixed with the air.

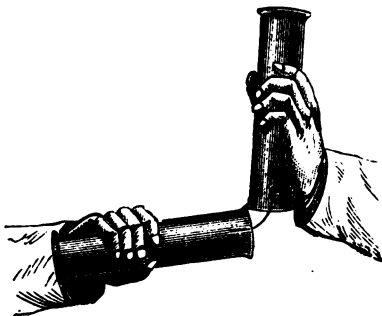


FIG. 5.

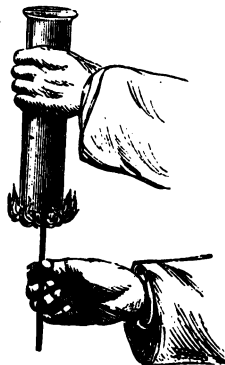
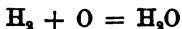


FIG. 6.

Jar 3.—Hold the jar mouth downwards (Fig. 6), and push up into it a lighted taper (a piece of taper tied on the end of a wire). The hydrogen burns where it comes in contact with the oxygen of the air, but the taper is put out. It may be relit at the mouth of the jar, and put out again several times.

Experiment.—Remove the delivery tube from the flask, and insert a small piece of glass tubing drawn out to a jet at one end, and cut slanting at the other (see Fig. 7). The air is all expelled, and the jet may therefore be lighted if the tube is inserted quickly. Hydrogen mixed with air *explodes* when a flame is applied to the mixture. *Never bring a lighted taper near a hydrogen apparatus till you are quite sure it is free from air.* The jet burns steadily with formation of water.



Hold over the flame a *dry* beaker; the water formed by the union (burning) of the hydrogen and oxygen rises as steam, and is condensed on the cold sides of the beaker.



FIG. 7.

Summary of properties. Hydrogen burns in air; is lighter than air; will not support combustion; when burnt in the air forms water.

The residue in the flask is a solution of zinc sulphate. Filter into an evaporating dish or beaker, and boil down to a small bulk. Allow to cool; when cold notice crystals of ZnSO_4 . If any zinc remain undissolved, wash it with water for future use.

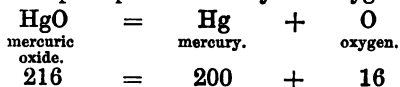


FIG. 8.

7. OXYGEN (O).

Preparation. (1) By heating mercuric oxide (HgO).

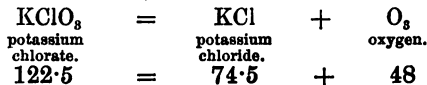
Place a little mercuric oxide in a dry test tube and heat. The mercuric oxide is split up into mercury and oxygen; thus,



The red powder becomes black, and little globules of mercury appear on the cool part of the tube. Keep the thumb loosely on the mouth of the tube while heating; and test for oxygen by lighting a match, blowing it out again, and inserting the spark, which still remains at the end of the match, into the tube. The match is rekindled.

(2) By heating potassium chlorate (KClO_3).

Heat gently a little potassium chlorate in a test tube; it melts, effervesces, and gives off oxygen. Test for oxygen as above. Potassium chloride remains.



Distinction between potassium chlorate and potassium chloride. When cold, add a little water to the residue in the tube used in the above experiment; the potassium chloride dissolves; add one or two drops of silver nitrate; a white curdy precipitate of silver chloride is formed (see tests for HCl , page 47).



Dissolve a crystal of potassium chlorate in distilled water, and add a drop or two of silver nitrate ; no precipitate is formed.

(3) By heating a mixture of potassium chlorate and manganese dioxide.

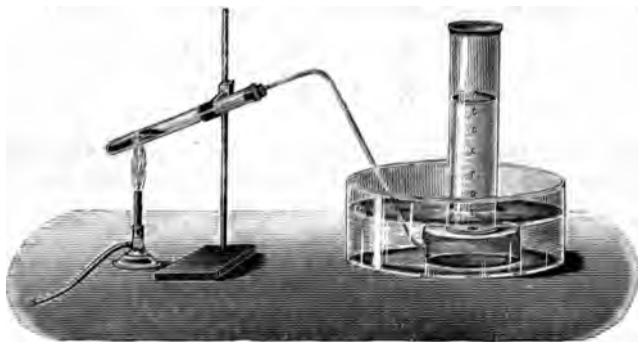
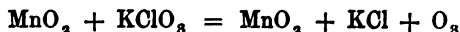


FIG. 9.

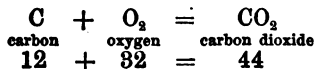
Fit up the apparatus similar to that shown in Fig. 9. Use a large test tube and fill it one-third full with potassium chlorate mixed with about one-fourth of its weight of manganese dioxide. This mixture liberates oxygen at a lower temperature and with greater rapidity than potassium chlorate alone. The oxygen is obtained from the potassium chlorate, the manganese dioxide remaining unaltered :—



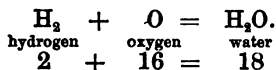
Apply heat, beginning at the upper surface of the substance and gradually moving the burner towards the bottom of the tube. Collect five jars of the gas, three at first and two more afterwards. When the oxygen ceases to be evolved remove the heat ; take the delivery tube out of the water, otherwise as the gas in the test tube cools and contracts, the cold water will be forced up the tube and will crack the apparatus.

Jar 1.—Fix a piece of taper to an iron wire ; light it and immerse it in oxygen. It burns much more brightly than in air. Take the taper out, blow it out, leave a spark, and again immerse it. The spark bursts into flame. No other gas behaves in the same way (except nitrous oxide).

The taper contains carbon and hydrogen; the carbon unites (burns) with the oxygen to form carbon dioxide,



the hydrogen unites (burns) with the oxygen to form water.

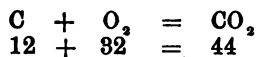


To prove the presence of carbon dioxide in the jar after burning the taper, add clear lime water and shake up; the lime water becomes milky (see page 17).

To prove the formation of water when a taper burns, take a perfectly dry beaker and invert it over the flame of a taper burning in the air. Notice the immediate appearance of condensed steam on the sides of the beaker.

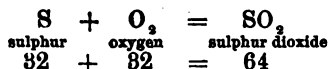
The result of the combustion of a taper is precisely the same whether it is burnt in air or oxygen; air is oxygen very much diluted with nitrogen (79 vols. of N to 21 vols. of O). The nitrogen takes no part whatever in the burning.

Jar 2.—Place a piece of charcoal in a deflagrating spoon, ignite it in the Bunsen flame, and immerse it in oxygen. It burns very brightly with formation of carbon dioxide gas.

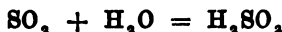


Remove the charcoal, add lime water and shake up; the lime water is turned milky.

Jar 3.—Place a small piece of sulphur in a deflagrating spoon, ignite in the Bunsen flame, and immerse it in oxygen. The sulphur burns with a beautiful blue flame, forming sulphur dioxide gas.

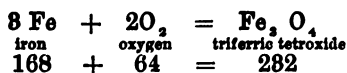


On addition of water to the jar this gas readily dissolves in the water, forming sulphurous acid (H_2SO_3).

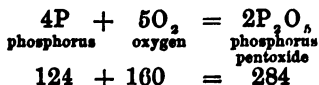


Pour into the jar a little blue litmus solution, it is turned red. Acids turn blue litmus red.

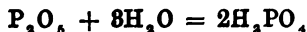
Jar 4.—Pour a little water in this jar, just sufficient to cover the bottom, to prevent the sparks from cracking the glass. Coil about 18 inches of very fine iron wire round a piece of glass tubing, so as to make a spiral. Fix one end of the wire in a cork to fit the jar, and attach to the other end a piece of wax taper. Set fire to the taper and immerse the whole in the oxygen. The heat from the burning taper enables the iron to unite with the oxygen; beautiful sparks of oxide of iron fly out in all directions.



Jar 5.—The teacher will cut off (under water) a piece of phosphorus, about the size of a small pea. Carefully dry it between blotting paper, and put it on a clean *dry* deflagrating spoon; set fire to the phosphorus by touching it with a hot wire and immerse in the oxygen. Combustion takes place with intense brilliancy; the bottle is filled with dense white clouds of solid phosphorus pentoxide (P_2O_5)



In a short time the white clouds disappear, being dissolved by the water in the jar with formation of phosphoric acid (H_3PO_4).

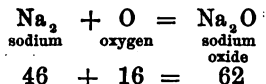


Add blue litmus solution, it is turned red.

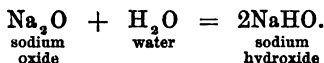
Wipe a gas jar perfectly dry and fill it with oxygen by downward displacement.

Jar 6.—See that your fingers are dry. Cut off a small piece of sodium about the size of a pea. Be very careful to keep the sodium away from water. Place the sodium in a dry and clean deflagrating

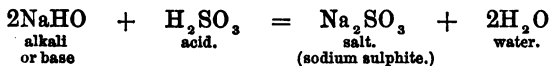
spoon ; ignite in the Bunsen flame and immerse in the dry jar of oxygen. It burns brilliantly, forming sodium oxide.



When cold, dissolve some of the white sodium oxide remaining in the spoon, in water in a beaker ; sodium hydroxide (NaHO) is formed :—



Place the beaker containing the solution of sodium hydroxide, side by side with the jar in which the sulphur was burned, and which now contains a solution of sulphurous acid. Into a clean test glass pour a little blue litmus solution ; add to it a little of the sulphurous acid ; the blue litmus is turned red. Now add drop by drop the sodium hydroxide solution ; a point is soon reached when a further addition of one drop of the sodium hydroxide will turn the solution blue. This is the *neutral* point. These two substances, sodium hydroxide and sulphurous acid, are examples of the two classes into which hydroxides are divided, viz. :—*Alkaline* or *basic* hydroxides, and *acid* hydroxides. They are opposite in properties, and when united they neutralize one another forming salts, thus :—



Summary. Substances burn more readily in oxygen than in air. Elementary bodies burn in oxygen to form oxides. The weight of the oxide formed is always greater than the weight of the substance taken before its union with the oxygen ; the weight of the oxide, in fact, is exactly equal to the weight of the substance taken plus the weight of the oxygen with which it united (see equations). Some oxides unite with water to form *alkaline* hydroxides ; other oxides unite with water to form *acid* hydroxides.

The residue in the oxygen apparatus is a mixture of potassium chloride and manganese dioxide. Add water to the test tube and warm ; the potassium chloride dissolves ; filter ; the manganese dioxide remains as a residue on the filter, the solution of potassium chloride passes through. Test a little of the solution for the presence of chloride with silver nitrate as before. Boil down the remainder of the solution in an evaporating dish to a small bulk ; *crystals of potassium chloride* separate out on cooling.

8. CHLORINE (Cl).

Preparation. By heating together manganese dioxide and strong hydrochloric acid.

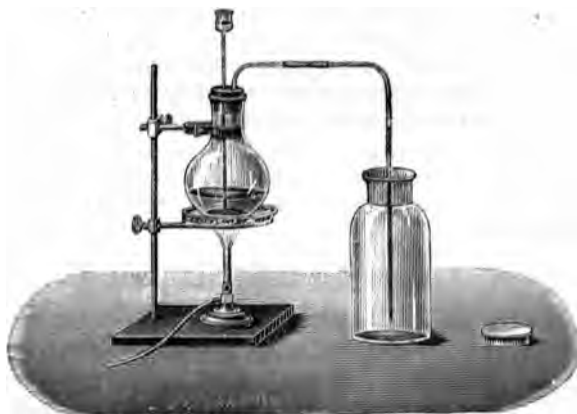
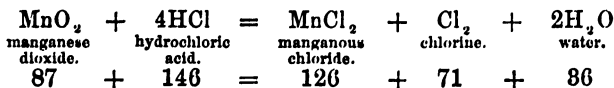


FIG. 10.

Fit up an apparatus as shown in Fig. 10.

The apparatus being ready, pour a little HCl into the flask and rinse it round, so as to wet the inside.

Now add about an ounce of manganese dioxide in the form of small lumps, free from powder, if possible, and pour on sufficient strong commercial HCl to cover the solid; shake until well mixed, fix the cork, and heat over a sand bath.

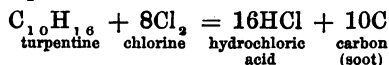


Chlorine, a yellowish green gas, is readily given off, and manganous chloride and water remain. As the gas is very injurious when breathed, it should be prepared in a draught closet. Chlorine cannot be collected over cold water, as water at 17°C dissolves about $2\frac{1}{2}$ times its volume of the gas; but, being a heavy gas, it is conveniently collected by downward displacement (density of $\text{Cl} = 3.5$, density of air = 1.42). See that your jars are dry. Collect five jars of the gas, three at first and two more afterwards; the jars are full when the yellowish green colour of the gas extends to the top of the jar, seen best by putting a sheet of white paper.

behind the jar. Cover the mouth of each jar with a greased glass plate. Let the chlorine also pass into a beaker containing cold distilled water. The water will become yellow; it possesses all the properties of the gas, and is termed *chlorine water*.

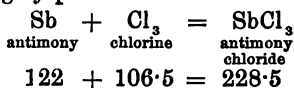
Jar 1.—Place a lighted taper in the jar. The taper continues burning, but with a dull, smoky flame. The hydrogen of the taper unites with the chlorine (hence the burning) to form HCl; the carbon of the taper is deposited as soot.

Jar 2.—Moisten a strip of filter paper, or a piece of tow, with turpentine, and immerse in the chlorine. Turpentine ($C_{10}H_{16}$) also consists of hydrogen and carbon; therefore HCl is formed, and soot again deposited.

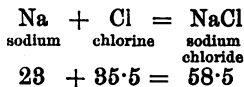


The paper or tow often takes fire.

Jar 3.—Take some finely-powdered antimony and shake a little from the end of a knife into the jar. Brilliant sparks appear, and chloride of antimony is formed. Put the glass plate on immediately, as these fumes are highly poisonous.



Jar 4 (perfectly dry).—Put a small piece of sodium on a *dry* clean deflagrating spoon; heat it in the Bunsen flame till melted, and immerse in the chlorine. The sodium unites (burns) with the chlorine, forming sodium chloride or common salt.

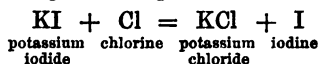


Jar 5.—Put a piece of coloured cloth, with one end moistened and the other end dry, into a jar of chlorine. The moistened part of the cloth is bleached, while the dry part remains unaltered. The chlorine unites with the hydrogen of the water liberating oxygen, which unites with the colouring matter to form a colourless compound.

The chlorine water. To a little litmus or indigo solution in a test glass, add chlorine water. The colour is destroyed.

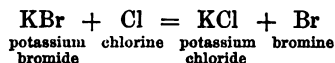
Experiment 1.—To a weak solution of potassium iodide (KI) add a little starch solution; then one or two drops of chlorine water; *the clear solution becomes deep blue in colour.*

The chlorine decomposed the potassium iodide, liberating iodine.



The liberated iodine unites with the starch, forming a blue compound.

Experiment 2.—To a solution of potassium bromide (KBr) add a drop of chlorine water; the clear solution becomes yellow, owing to the liberation of bromine.

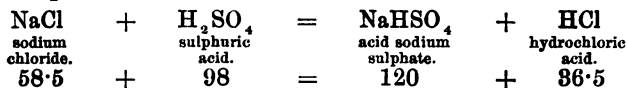


Summary of properties. Chlorine is a yellowish green gas, heavier than air; soluble in water; does not burn; a taper burns in it with a smoky flame; metals burn in it to form chlorides; it is a powerful bleaching and oxidising agent; displaces iodine and bromine from their salts.

9. HYDROCHLORIC ACID (HCl).

Preparation. By the action of strong sulphuric acid on sodium chloride.

Fit up an apparatus similar to that for chlorine, Fig. 10. Place about half an ounce of common salt in the flask (rock salt broken into small lumps is better). Add some strong sulphuric acid very gradually by the thistle funnel tube, the end of which should be below the solid salt. The exact proportion of acid to salt is seen by the equation—



Hydrochloric acid gas is liberated, and acid sodium sulphate remains.

The gas comes off at first without heating, and, if common loose salt is used, with much frothing. As soon as it slackens, heat gently on a sand bath. Owing to the great solubility of hydrochloric acid gas, it cannot be collected over water, but is collected by downward displacement. (Density of

$$\text{HCl} = \frac{1 + 35.5}{2} = 18.25; \text{ density of air, } 14.42.)$$

Thoroughly dry your gas jars, and collect two jars of the gas. When fumes proceed from the mouth of the jar, it is full; cover

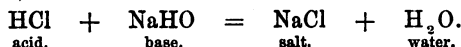
each with a glass plate. Make a solution of HCl by passing the gas into a beaker containing a small quantity of water. The gas is quickly absorbed by water, and care must be taken that the water does not flow up the tube; in such a case increase the heat, and, if necessary, remove the beaker.

Jar 1.—Invert the jar, and place it mouth downwards in water coloured with blue litmus; remove the glass plate. The HCl gas is absorbed by the water, a vacuum is thus formed, and the water is forced in by the air and fills the jar. The blue colour of the water is changed to red, showing that HCl is an acid gas.

Jar 2.—Wet a strip of filter paper with ammonia solution, and hold it in the jar; dense white clouds of ammonium chloride are formed.



The solution.—Add a little blue litmus solution, it is at once turned red. Now add some sodium hydrate solution and stir with a glass rod; when sufficient alkali (NaHO) has been added to neutralize all the acid (HCl), the addition of another drop of sodium hydrate will turn the solution distinctly blue. Carefully obtain the neutral point; taste the solution; a salt, namely sodium chloride (common salt), has been formed.



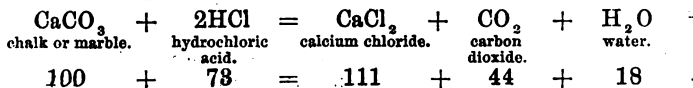
Evaporate to a small bulk and obtain crystals of sodium chloride.

Summary of properties. Hydrochloric acid gas is heavier than air, is very soluble in water, reddens blue litmus, forms salts called chlorides when added to bases.

10. CARBON DIOXIDE (CO₂).

Preparation. (1). By the action of acids on carbonates.
(2). By burning carbon (charcoal) in air or oxygen.

Fit up an apparatus similar to that for hydrogen, Fig. 4. Place in the flask some pieces of marble or chalk (calcium carbonate); pour on water till the pieces are covered, and add hydrochloric acid through the thistle funnel tube; great effervescence occurs, due to the escape of carbon dioxide gas from the carbonate, thus—



Collect two jars of the gas over water, and cover each with a glass plate; then pass the gas into a beaker containing lime water (Fig. 11); the lime water becomes milky owing to the formation of insoluble calcium carbonate (chalk), the little solid particles of which float about in the water.

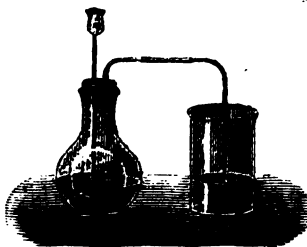
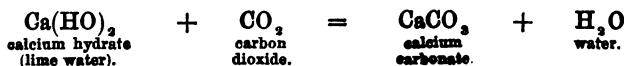
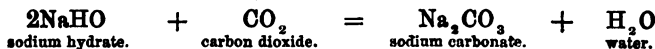


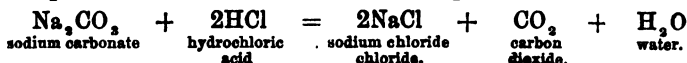
FIG. 11.



Again pass the gas for a few minutes into a solution of sodium hydrate contained in a test glass or beaker; sodium carbonate is formed, but unlike the calcium carbonate it is soluble in water, hence no milky or precipitate appears.



To show that a carbonate has actually been formed, add a few drops of HCl, and notice the effervescence which occurs, due to the escape of CO_2 from the sodium carbonate present.



Jar 1.—Introduce a lighted taper; the gas does not burn, and the taper is immediately put out. Pour in a little blue litmus solution and shake up. The litmus becomes deep red in colour; carbon dioxide added to water becomes carbonic acid.



Boil this red solution in a beaker, the carbonic acid (H_2CO_3) is decomposed, CO_2 is driven off, and the red solution again becomes blue.

Take a clean, empty jar (containing air only), pour into it a little lime water and shake up; the lime water does not become milky.

Jar 2.—Pour the carbon dioxide from this jar into the clean jar just mentioned, in the same way as you would pour water; the carbon dioxide, being heavier than air, falls into the lower jar, taking the place of the air. Again shake up the lime water; it is at once turned milky.

Summary of properties. Carbon dioxide turns lime water milky; forms carbonates when passed into solutions of hydrates; does not burn; puts out a lighted taper; forms an acid with water; is heavier than air.

The apparatus.—Pour off part of the solution of calcium chloride on a filter; test the clear filtrate for calcium and chlorine by 25, Ca, 4; and 26, HCl 1.* Throw away the remaining liquid, and wash the marble or chalk with water for future use.

11. NITROGEN (N).

Preparation. By removing the oxygen from a confined volume of air by the aid of phosphorus.



FIG. 12.

Dry a small piece of phosphorus between blotting paper (see oxygen, jar 5), place it in a little dry porcelain dish or cover, and allow the dish to float on water in a pneumatic trough. Ignite the phosphorus with a hot wire, and cover it at once with a bell jar, tightly corked. At first a few bubbles of air escape, owing to expansion through the heat of combustion. Dense white clouds of phosphorus pentoxide (P_2O_5) are formed, which are soon dissolved in the water to form phosphoric acid (H_3PO_4), and the water rises about one-fifth up the jar to take the place of the oxygen which has united with the phosphorus. The gas which remains is nitrogen. When the white clouds have settled, remove the cork and put a lighted taper in the jar. It is at once extinguished. Nitrogen will not burn nor support combustion.

* 26, Ca, 4, means paragraph 26 (Table of Reactions), calcium column, test number 4.

Add blue litmus solution to the water in the trough. It is turned red by the phosphoric acid.

Summary. Air consists of oxygen and nitrogen. The gases are mixed together in the proportion by volume of 79 of nitrogen to 21 of oxygen (nearly 4 to 1). Nitrogen will not burn, nor allow a taper to burn in it.

NOTE.—Do not attempt to ignite the small particle of phosphorus which may remain unburned, as it will probably be wet, and would fly about. It should be put into a bottle containing water.

12. AMMONIA (NH_3).

Preparation. By the addition of one of the alkaline hydrates—*e.g.* KHO , NaHO , $\text{Ca}(\text{HO})_2$ —to any salt of ammonium; usually by heating together lime $\text{Ca}(\text{HO})_2$ and ammonium chloride.

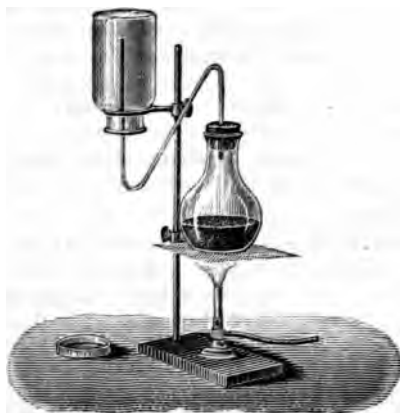
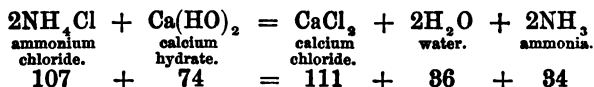


FIG. 13.

Fit up an apparatus as shown in Fig. 13.

Powder some ammonium chloride in a mortar, mix with this about the same weight of slaked lime, and put the mixture into the flask. Add a little water, and stir so as to make a thick paste. Fix the cork and tube, let the gas jars be quite dry, and heat gently

over a sand bath or wire gauze. Ammonia is rapidly given off; thus,



Calcium chloride and water remain in the flask.

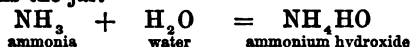
Owing to its extreme solubility, ammonia cannot be collected over water, but being lighter than air

$$(\text{density of NH}_3 = \frac{14+3}{2} = 8.5; \text{ density of air} = 14.42)$$

it is readily collected by upward displacement. To find when the jar is full, hold a piece of moistened red litmus paper just above the mouth of the jar on the outside; when the ammonia overflows, the paper is turned blue.

Collect two jars of the gas, and cover each with a glass plate; then pass the gas for a few minutes into a little distilled water in a beaker. See that the water does not flow back into the apparatus.

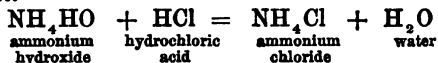
Jar 1.—Invert it in a vessel of water coloured with red litmus solution and remove the glass plate; the ammonia is absorbed by the water, a vacuum is thus formed, and the water is forced in by the air and fills the jar.



The red colour of the water is changed to blue, showing that ammonia has an alkaline reaction; alkalies turn red litmus blue.

Jar 2.—Moisten a strip of paper with HCl, and immerse it in the jar; white clouds of ammonium chloride (NH_4Cl) are formed.

The solution. Take a little blue litmus solution in a test tube, and add one drop of HCl, the blue colour changes to a bright red; now add one or two drops of the ammonia solution, the colour is again changed to blue. Ammonium hydroxide, like sodium hydroxide, neutralizes acids forming salts. Neutralize a small quantity of HCl with your ammonia solution; use litmus paper to detect the neutral point.



Evaporate carefully to dryness in an evaporating dish and observe the white solid ammonium chloride obtained. Continue the heat strongly in the draught cupboard; white clouds arise and the ammonium chloride completely disappears (volatilizes).

Summary of properties. Ammonia is a very light gas ; is very soluble in water, forming ammonium hydroxide ; turns red litmus blue ; neutralizes acids, forming salts of ammonium ; salts of ammonium are volatile.

13.—NITRIC ACID (HNO_3).

Preparation. By the addition of strong sulphuric acid to potassium nitrate (nitre).

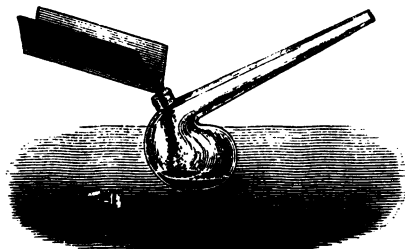
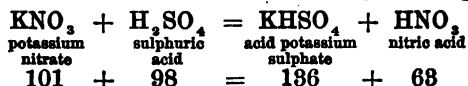


FIG. 14.

Fit up an apparatus similar to that for distilling in Fig. 2. Remove the stopper and place in the retort an ounce of potassium nitrate (see Fig. 14), then add carefully through a funnel (placed in the hole for the stopper) sufficient strong sulphuric acid to cover the salt, see that none of the nitre or sulphuric acid gets into the neck of the retort. Replace the stopper, and when ready apply a gentle heat to the bulb of the retort through a wire gauge. Allow a small stream of cold water to flow over the flask. The nitric acid readily distils over and is condensed in the flask as a yellow liquid.



The red fumes which will be observed are due to the escape of lower oxides of nitrogen.

When the boiling has ceased in the retort, and no more red fumes are seen, remove the burner, disconnect the apparatus, remove the stopper, and while the substance is still liquid pour it into a dry evaporating dish. The substance solidifies ; it is acid potassium sulphate. (See equation.) Be very careful in pouring from a retort, to hold the stem in an upright position.

The nitric acid. Test separate small portions of the liquid in the flask by the nitric acid tests. (26, Nitric acid).

14. NITROUS OXIDE (N₂O).

Preparation. By heating ammonium nitrate.

Prepare an apparatus as shown in sketch ; use a dry flask.

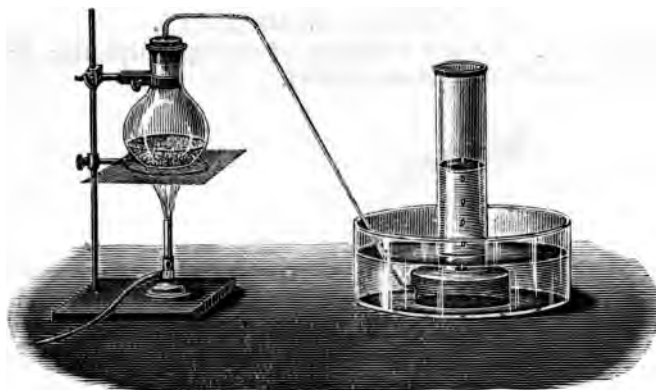
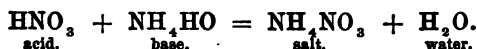


FIG. 15.

Preparation of the ammonium nitrate.—Pour some dilute nitric acid into an evaporating basin ; now add carefully, with continued stirring, dilute ammonia solution till the mixture is neutral to test paper. The nitric acid and ammonium hydroxide react upon one another, forming ammonium nitrate and water, thus :—

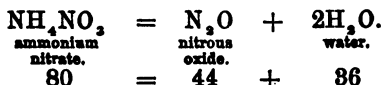


Evaporate nearly to dryness, then warm till dry on a sand bath. When cold break up the salt into small pieces and place in the flask.

The pneumatic trough should be supplied with *warm* water, as nitrous oxide is less soluble in warm than in cold water.

Apply heat carefully to the flask through a wire gauze (a strong heat may break the flask) ; at first air is driven off owing to its expansion by the heat ; the salt melts and froths up ; remove the burner till the frothing subsides ; *use a small flame*, and proceed to collect three jars of the gas. When the jars are full, remove the delivery tube from the water immediately, (many accidents occur through neglect of this precaution).

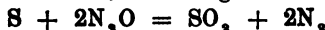
The ammonium nitrate is split up into nitrous oxide and water, thus:—



Jar 1.—Immerse a lighted taper; it burns in the nitrous oxide much more brightly than in air. A glowing chip is also rekindled, as in oxygen. The nitrous oxide is split up into nitrogen and oxygen, and this mixture contains a larger proportion of oxygen than is contained in air, hence the increased vigour of combustion. ($\text{N}_2\text{O} = 1$ vol. of O to 2 vols. of N; air = 1 vol. of O to about 4 vols. of N.)

Jar 2.—Ignite slightly a small piece of sulphur in a deflagrating spoon, and immerse it in the jar; it ceases to burn; take it out, replace the glass plate, ignite the sulphur strongly, and again immerse it. The sulphur now burns much more brightly than in air.

Sulphur dioxide is formed, and nitrogen remains mixed with it.



Substances will not burn in nitrous oxide until it has first been split up into nitrogen and oxygen, and the temperature of the feebly burning sulphur was not sufficiently high to decompose the gas.

To distinguish between nitrous oxide and oxygen.

Prepare a jar of oxygen.

Jar 3.—Invert this jar (of N_2O) in a vessel of cold water, and place by its side, also inverted in the water, the jar of oxygen. Notice (1) that the water rises slowly in the jar of nitrous oxide, because this gas is soluble in cold water, while the water does not rise in the jar of oxygen. Now prepare, in a small flask fitted with a cork and delivery tube, nitric oxide gas—for preparation see page 24). When pure, allow this gas to bubble first into the jar of oxygen. Notice (2) that deep red fumes are formed, due to the union of the nitric oxide and oxygen. These fumes are very soluble, and the water at once rises in the jar. Now pass the nitric oxide (NO) into the jar of nitrous oxide (N_2O); no red fumes appear, and the water does not rise in the jar.

Nitrous Oxide.

1. Readily soluble in water.
2. No action with nitric oxide.

Oxygen.

1. Not readily soluble in water.
2. Forms red soluble fumes with nitric oxide.

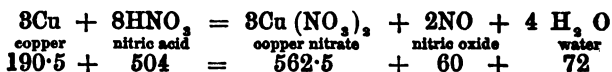
Summary of properties.—Nitrous oxide is colourless, is soluble in cold water; a glowing chip is rekindled in it as in oxygen; feebly burning sulphur is extinguished in the gas, but strongly ignited sulphur burns brilliantly.

15. NITRIC OXIDE (NO or N_2O).

Preparation. By the action of nitric acid on copper.

Fit up an apparatus similar to that for hydrogen, Fig. 4.

Place some copper turnings in the flask, cover them with water and replace the cork. Add gradually by the thistle funnel tube strong nitric acid: nitric oxide gas is liberated, thus



A solution of copper nitrate remains. At first the flask is filled with deep red fumes (N_2O_3 and N_2O_4) caused by the union of the nitric oxide with the oxygen of the air contained in the flask. As soon as the air is all expelled, the gas becomes colourless, and the pure nitric oxide may now be collected. Collect three jars of the gas and keep the mouth of each under water.

Jar 1.—Uncover the mouth of the jar and expose the gas to the air. The nitric oxide unites with the oxygen of the air forming dense red fumes of nitrogen trioxide (N_2O_3) and nitrogen tetroxide (N_2O_4). Now invert the jar in a vessel of water; the red gases are soluble and the water rises rapidly in the jar. (See also addition of nitric oxide to oxygen, page 28.)

Jar 2.—Remove the glass plate and quickly insert a lighted taper; it is at once extinguished. Replace the plate; heat a small piece of sulphur on a spoon in the Bunsen flame and immerse in the gas; the burning sulphur is extinguished.

Jar 3.—Take a piece of phosphorus about the size of a small pea; dry it carefully between blotting paper; place it on a clean, dry deflagrating spoon; touch it with a hot wire and immerse quickly in the jar. The phosphorus ceases to burn. Remove the phosphorus and replace the glass plate; rekindle the phosphorus in the Bunsen flame till it burns brightly, and again immerse it in the gas. The phosphorus now burns with much more brilliancy than in air; phosphorus pentoxide (P_2O_5) is formed and nitrogen remains.

Summary of properties. Nitric oxide is colourless ; it unites with oxygen to form red gases (higher oxides of nitrogen). It is more difficult to decompose nitric oxide than nitrous oxide, hence the former (NO) does not so readily support combustion as the latter (N_2O). The temperatures of a lighted taper, burning sulphur, and feebly burning phosphorus are not sufficiently high to decompose nitric oxide ; hence these bodies do not burn in it. Strongly burning phosphorus, however, decomposes the gas into nitrogen and oxygen and burns brilliantly in the liberated oxygen.

The apparatus.—Pour a small portion of the solution of copper nitrate on a filter ; test the clear filtrate for copper and nitric acid (25, Cu 4, and 26, nitric acid). Pour away the remainder. Wash with water any copper remaining undissolved and save it for future use.

16. SULPHURETTED HYDROGEN (H_2S).

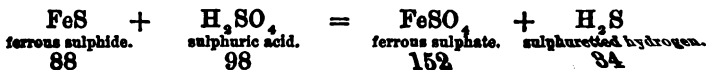
Preparation. By the action of sulphuric acid (or of HCl) on ferrous sulphide.



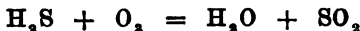
FIG. 16.

Fit up the apparatus as in Fig. 16.

Place in the flask with the thistle funnel tube some small pieces of ferrous sulphide (FeS) ; cover the sulphide with water, and replace the cork. The second flask is a wash bottle, and is half filled with water to absorb impurities. Add through the thistle funnel tube a little strong sulphuric acid ; ferrous sulphate is formed, and sulphuretted hydrogen gas is liberated. It is at once recognised by its characteristic smell.



A jar of H_2S .—Collect one jar of the gas by downward displacement (density of $H_2S = \frac{2 + 32}{2} = 17$; density of air = 14.42) in a draught cupboard. When full ignite the gas in the jar with a taper; it burns with a bluish flame; the hydrogen of the H_2S burns to form water, and the sulphur to form sulphur dioxide; part of the sulphur is deposited on the sides of the jar.



A solution of H_2S .—Pass the gas into a flask containing cold, recently boiled, distilled water until the water smells strongly of the gas. Sulphuretted hydrogen is soluble in water, and we now have a solution of the gas.

This solution spoils if exposed to the air for some time; the hydrogen of the H_2S unites with the oxygen of the air, and the sulphur is deposited in the water, giving the solution a milky appearance.

The apparatus.—Pour off the solution containing ferrous sulphate dissolved, and probably an excess of sulphuric acid; wash the remaining sulphide with water for future use.

Solutions for testing.—Take separate small quantities in test tubes of the following solutions: add first a few drops of HCl to each, except to the lead solution (HCl would give a white precipitate of lead chloride), then a little of your H_2S solution. Notice the results, and proceed as required in the Table.

<i>Solutions.</i>	<i>Result of adding H_2S to acid (HCl) solution</i>	
Copper sulphate, $CuSO_4$ Lead nitrate, $Pb(NO_3)_2$ Arsenious acid, H_3AsO_3 Antimony chloride, $SbCl_3$	Black ppte. of CuS Black ppte. of PbS Yellow ppte. of As_2S_3 Orange ppte. of Sb_2S_3	
Zinc sulphate, $ZnSO_4$ Ferrous sulphate, $FeSO_4$ Magnesium sulphate, $MgSO_4$	no ppte. no ppte. no ppte.	<i>Make the Zn, Fe and Mg solutions alkaline with NH_4OH</i> White ppte. of ZnS Black ppte. of FeS no ppte.

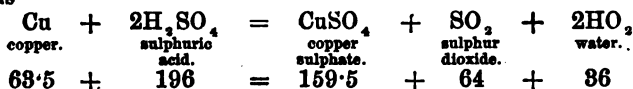
For the Equations see Part II., Question 289.

17. SULPHUR DIOXIDE (SO_2).

Preparation. (1) By burning sulphur in oxygen. (2) By the action of strong sulphuric acid on copper.

Fit up an apparatus similar to that used for chlorine in Fig. 10.

Place in the flask about half an ounce of copper turnings; pour through the thistle funnel tube strong sulphuric acid till the copper is covered, and heat the flask carefully over a sand bath. Sulphur dioxide gas is given off, and a solution of copper sulphate remains; thus—



The gas being very soluble in water, and heavier than air

$$(\text{density of } \text{SO}_2 = \frac{32 + 32}{2} = 32; \text{ density of air} = 14.42)$$

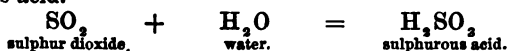
it is collected by downward displacement.

Collect two jars of the gas; use dry jars, and when full cover with a glass plate.

Pass the SO_2 for a few minutes into a small bottle three-fourths full of distilled water; replace the stopper in the bottle.

Now pass the gas into small portions of the following solutions: dilute blue litmus, potassium permanganate, nitric acid, potassium chromate.

The *litmus solution* is turned red, showing that sulphur dioxide solution is acid.



The *potassium permanganate* ($\text{K}_2\text{Mn}_2\text{O}_8$) loses its brilliant purple-red colour and becomes colourless. Here the sulphurous acid (H_2SO_3) *reduced* (robbed of oxygen) the permanganate; the sulphurous acid becoming oxidised to sulphuric acid (H_2SO_4).

The *potassium chromate* (K_2CrO_4) changes colour from yellow to green owing to the reducing action of sulphurous acid.

Jar 1. Immerse a burning taper, the flame is at once put out, and the gas does not burn. Quickly replace the glass plate, invert the jar and place it in water. The sulphur dioxide is dissolved in the water; sulphurous acid is formed, and the water rises rapidly in the jar.

Jar 2. Place in this jar a moistened bunch of flowers ; they will be bleached, but the colour may be restored if dipped into a weak solution of acid. Bodies such as straw, silk, wool, are bleached by exposing them to the action of sulphur dioxide ; chlorine injures such fabrics.

The solution of SO_2 . With this we can show that sulphurous acid reduces nitric acid ; or, in other words, that nitric acid oxidises sulphurous acid : we can also compare the result of adding hydrochloric acid and barium chloride (BaCl_2) to solutions of sulphurous and sulphuric acids respectively.

Pour part of the solution into two clean test glasses. To one add barium chloride ; a white precipitate of barium sulphite (BaSO_3) is produced, which dissolves, however, on addition of HCl .

To the other test glass add a few drops of nitric acid ; the sulphurous acid is now converted into sulphuric acid. Add barium chloride ; a dense white precipitate of barium sulphate (BaSO_4) falls, which is not dissolved on addition of HCl (see equations, page 82).

The *apparatus*. Pour a little of the copper sulphate solution into a test tube, and test for copper and sulphuric acid by special tests (25, Cu 4, and 26, sulphuric acid). Pour away the remainder, wash any remaining copper with water, and set aside for future use.

Summary of properties. Sulphur dioxide is colourless ; heavier than air ; does not burn nor support combustion ; soluble in water ; forms with water sulphurous acid ; bleaches vegetable colouring matters ; is a powerful reducing agent.

18. COMBUSTION.

When elements unite together to form a new compound, heat is given off, and sometimes light. Bodies which unite together vigorously are said to *burn* or *undergo combustion*. Burning or combustion is therefore the result of the union of two or more substances. One substance cannot burn by itself.

For example, we say hydrogen will burn, but it is understood that we mean hydrogen will burn *in the air*. The hydrogen unites (burns) with the oxygen of the air to form water. It should be remembered that in the flame produced by the burning of hydrogen in air, for every 2 parts by weight of hydrogen consumed there are 16 parts by weight of oxygen from the air consumed also, forming *18 parts by weight of water*.

The young student might suppose that when a candle burns in a room it is only the candle, and nothing else, that is affected by the burning, forgetting that there can be no burning without combination. Without the oxygen of the air the candle would go out.

Experiment 1.—Immerse a small burning candle in a bottle stopped with a cork. The oxygen of the air in the bottle is soon all consumed, or combined with the hydrogen and carbon of the candle, and the candle goes out.

We say oxygen will not burn, but we mean that oxygen will not burn in the air. The reason is because there is nothing in the air with which the oxygen will readily combine.

The following experiment, however, shows that even air may be made to burn :—

Experiment.—Take a broad glass tube open at both ends—a common oil-lamp chimney is best. Fit to one end of the tube a cork bored with two holes. Through one of the holes pass a short piece of metal tubing about $\frac{1}{4}$ in. diameter (a piece of sheet tin will do, rolled into a tube), and through the other hole put a piece of glass tubing to which to attach a piece of india-rubber tubing from the gas nozzle.

Attach the apparatus to a stand, as seen in Fig. 17.

Place a piece of fine wire gauze over the top of the large tube.

Pass coal gas into the apparatus, and light it at the top of the gauze. The cylinder is now full of coal gas. Pass a lighted taper up the metal tube. A flame appears on the top of the metal tube, and we now see (i) the *air burning* in coal gas inside the large tube, and (ii) coal gas burning in the air outside the apparatus.

In the above experiment it is the oxygen of the air only which burns, the nitrogen is quite passive, and takes no part whatever in the combustion.

Experiments to show that combination takes place between the gases at the point of contact only.

Experiment 1.—Spread a sheet of paper on a smooth surface and press the Bunsen flame for an instant on the paper. (Fig. 18.)

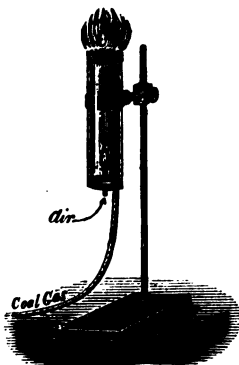


FIG. 17.

The exact shape of the flame will be burnt on the paper, while the space between the burnt marks is quite unaffected.



FIG. 18.

Experiment 2.—Place a piece of wire gauze upon the flame of a Bunsen burner. A red-hot ring will appear, corresponding with the outside edges of the flame.

Experiment 3.—Place a piece of platinum wire across the Bunsen flame. The wire becomes red hot at the two points where it enters and leaves the flame.

Each of the above three experiments shows that the heat of the flame is on its outside edges; it is here where chemical combination is taking place between the coal gas and the oxygen of the air.

19. THE CANDLE FLAME.

The flame of a candle consists of three principal parts:—

- (1) **The dark central zone**, containing the gaseous products of the tallow, chiefly composed of compounds of carbon and hydrogen in an unburnt state.
- (2) **The luminous (or light-producing) zone**, surrounding the central zone. Here, owing to an insufficient supply of oxygen, the gases are only partially burnt; the hydrogen burns chiefly, whilst the carbon remains in a highly heated, glowing state, causing the luminous appearance.
- (3) **The outer zone.** Here there is an abundant supply of oxygen, and combustion is complete. This portion of the flame is very hot, the greatest heat being concentrated at *a*.



FIG. 19.

Oxidising flame. When bodies are strongly heated in contact with oxygen, oxides are formed (see experiments with oxygen). Now, at the outer part of the candle flame we have both a high temperature and an unlimited supply of oxygen. The outer zone of the candle (8) is therefore known as the *oxidising flame*.

Reducing flame. On the other hand, when oxides are placed in the *luminous* zone of the candle flame they have a tendency to yield up their oxygen; in other words, to become *reduced*, the oxygen being withdrawn from them by the heated carbon and hydro-carbons. Hence the luminous zone is known as the *reducing flame*.

20. THE BUNSEN BURNER.

We have seen that the *luminous* appearance of the candle flame is due to an insufficient supply of oxygen to consume the heated hydrocarbon compounds. Now, if by some means we could introduce a supply of oxygen at the centre of the flame, you would expect the luminous zone to disappear. This is accomplished by the simple arrangement known as the "Bunsen burner" (see Fig. 20).

The coal gas enters, as shown in the sketch, and passes up the large tube; at the same time, by means of the holes at the bottom of the tube, the air flows in and mixes with the coal gas, thereby supplying sufficient oxygen to destroy the luminosity of the flame, and to greatly increase its temperature.

If the air-holes be closed, the supply of oxygen to the centre is cut off, and the flame resumes its ordinary luminous appearance.

Experiment.—Hold a cold porcelain dish for a moment over the luminous flame; there will be a deposit of soot. The blue non-luminous flame gives no deposit of soot.

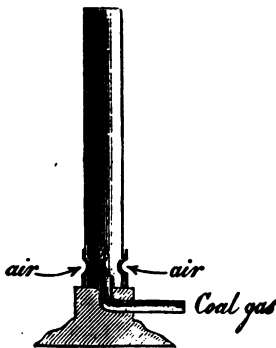


FIG. 20.

21. THE BUNSEN FLAME.

The Bunsen flame consists of two principal parts—(i) the inner cone of unburnt gas (ii) the outer mantle, formed by the burning mixture of coal-gas and air.

If the air-holes are partly closed, a luminous point appears at the top of the inner cone (*f*)—Fig. 21.

The following is a further description of various parts of the flame :—

- a.* The point of lowest temperature. Here volatile mixtures are heated to obtain flame colorations.
- b.* The hottest part of the flame.
- c.* The lower oxidising flame.
- d.* The upper oxidising flame ; temperature not so high as *c*.
- e.* The lower reducing flame.
- f.* The upper reducing flame ; possesses greater reducing power than *e*.

Many bodies may be detected by the characteristic colour which they impart to a non-luminous flame. The chlorides yield the most distinct colours, hence the substance to be tested is first moistened with HCl.

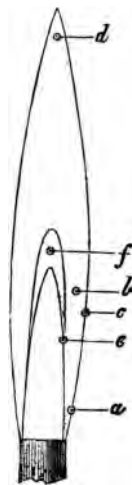


FIG. 21.

The following substances give characteristic flame colorations :—

Green.	Crimson.	Orange-red.	Violet.	Yellow.	Blue.
Barium. Copper.	Strontium.	Calcium.	Potassium.	Sodium.	Lead. Copper Chloride.

Clean your platinum wire by holding the end of the wire in the Bunsen flame till the yellow (sodium) colour disappears ; occasionally dipping in a little dilute HCl will assist to clean it. The wire should be bent into a loop at the end by twisting it round the sharpened point of a lead pencil. The other end of the wire should be fused into a piece of glass rod or tubing. Take a small particle of a salt of each of the above flame-colouring substances ; test each separately on the loop of the wire by holding it in the lower outer edge of the Bunsen flame. Notice the characteristic colours.

22. USE OF THE BLOWPIPE.

It is essential that the student should be able to keep up a steady, continuous current of air through the blowpipe without getting out of breath. Practise at first by closing the lips, extending the cheeks, and breathing through the nose. The steady blast is maintained by the cheeks and not by the lungs.



FIG. 22. Reducing Flame.



FIG. 23. Oxidising Flame.

Close the holes at the bottom of the Bunsen burner and obtain a luminous flame, or insert the tube as shown in Figs. 22, 23. Bring the nozzle of the blowpipe just to the outside of the flame and a little above the edge of the burner; blow gently and you will obtain a partly luminous flame. Fig. 22. This is the *reducing* flame.

Now hold the nozzle of the blowpipe a little further into the flame and blow more strongly; the flame is now blue and non-luminous. Fig. 23. This is the *oxidising flame*, the hottest part of which is at the point.

23. EXERCISES WITH THE BLOWPIPE.

Reduction on Charcoal.

Experiment 1. Take a piece of charcoal and with the tip of a penknife blade scoop out a little round cavity in the charcoal; place in the cavity a small heap of litharge (PbO), and direct upon it the reducing flame of the blowpipe. After a little time a bright metallic globule of lead appears.

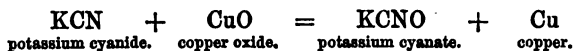
Take off the bead with your penknife, and let it fall into a clean mortar. Press it with a pestle, and notice that it is *malleable*, that is, may be flattened. Lay the flattened bead on a piece of clean paper and draw it along; it makes a mark on the paper.

Dissolve the bead in nitric acid in a test tube, and add H_2SO_4 ; a white precipitate of lead sulphate is obtained.

Silver compounds also yield a malleable bead of metallic silver, but this bead may be distinguished from lead since it makes no mark when drawn along the paper, and when dissolved in nitric acid gives no precipitate on addition of H_2SO_4 .

Experiment 2. Mix together a small quantity of litharge (PbO) and sodium bicarbonate (NaHCO_3) and heat in the reducing blowpipe flame. The sodium bicarbonate assists the reduction, and again the bright metallic bead of lead appears.

Some substances are reduced with greater difficulty than others, and frequently, to secure complete reduction, potassium cyanide (KCN) is added. Potassium cyanide is a more powerful reducing agent than sodium bicarbonate, the cyanide becoming itself oxidised to cyanate, thus—



Remember that potassium cyanide is *very poisonous*.

Experiment 3. Take a little copper oxide (CuO), mix it with some powdered KCN and NaHCO_3 and fuse on charcoal before the *reducing blowpipe* flame. Red particles of copper are obtained.

Some substances, on being reduced on charcoal, do not yield a metallic bead, but separate in small metallic particles, as in the following example :—

Experiment 4. Powder a little ferrous sulphate, mix it with NaHCO_3 , and fuse in the reducing blowpipe flame. Metallic particles of iron are obtained, which may be collected thus :—Scrape the fused mass from the surface of the charcoal into a clean mortar, powder the mass and add water. On stirring, the particles of iron sink, while the charcoal tends to float. Pour off the water and suspended charcoal, and repeat the washing ; stir gently with a magnetised knife blade ; the finely divided iron adheres to the blade, and may be wiped off upon filter paper.

Oxidation.

Experiment 1. Place a small piece of lead—the bead obtained in the above reduction experiment—and heat before the oxidising flame of the blowpipe. A yellow incrustation of litharge (PbO) is formed on the charcoal. By the aid of the blowpipe flame and the oxygen of the air, the lead bead has been oxidised back again to litharge.

The Borax Bead.

Borax ($\text{Na}_2\text{B}_4\text{O}_7, 10\text{H}_2\text{O}$) when heated swells up, owing to escape of water, and then melts down to a clear transparent mass.

Melted borax possesses the property of dissolving many metallic oxides, in some cases producing characteristic colours.

Experiment 1. Make a loop on the end of your platinum wire ; make it hot, and dip it in powdered borax. The borax clings to the wire, swells up, and on heating in the blowpipe flame, forms a transparent bead in the loop.

Detach the smallest particle of a cobalt compound on clean paper ; heat the bead, and touch the particle with it. Now heat again before the blowpipe. The cobalt has coloured the bead a beautiful deep blue. If too much cobalt is taken up in the bead, the blue will be so intense as to appear black.

Experiment 2. Clean your wire ; make another borax bead, and attach to it a trace of a copper compound ; heat in the oxidising flame ; a bead is obtained which is green whilst hot, blue on cooling. Now add to the copper bead a particle of any tin salt, and heat in the reducing flame, the colour of the bead changes to red owing to the formation of Cu_2O .

To make a bulb tube.

Take a piece of glass tubing about $\frac{1}{4}$ in. bore and about 6 inches long. This will make two tubes. Draw it out at its middle point (a) in the Bunsen flame, cut off the thin thread of glass with a file and seal up the end (b).

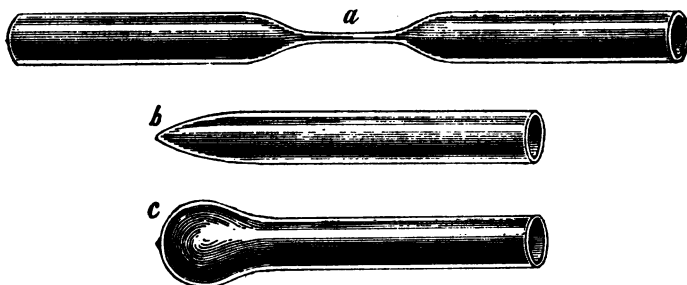


FIG. 24.

Blow a small bulb (c) by heating the end (b) in the Herepath blowpipe till the glass is red hot and softens ; blow carefully into the tube, turning it round at the same time. Skill in blowing bulbs is acquired only by practice. Cut off to the required length.

Experiment 1. Take a little mercurous chloride, and mix it with powdered NaHCO_3 . Introduce it into the bulb by placing it on a long narrow piece of stiff paper which has been folded lengthways. Heat the bulb, at first gently, afterwards more strongly ; a grey deposit of metallic mercury appears on the side of the tube ; scrape the deposit together with a strip of wood, little globules of mercury appear.

24. THE TABLE OF REACTIONS.

Solutions of the following salts are required :—Silver nitrate, lead nitrate, mercurous nitrate, mercuric nitrate or chloride, copper sulphate, ferrous sulphate, ferric chloride, calcium chloride ; also, a selection of solids to illustrate the dry-way tests. A test-tube stand and about a dozen test tubes will also be necessary.

The Table should be used as follows :—

Put a small quantity of each of the above solutions into separate test tubes, arrange them side by side and add to each, say first, HCl. Notice and compare the action of this reagent on the various solutions.

Afterwards, to fresh solutions, proceed in the same way to add H_2S , and so on with all the reagents in the left-hand column.

Compare also the behaviour of the solid substances with the dry-way tests.

The contraction ppte. = precipitate.

25. REACTIONS OF THE METALS—SILVER,

	SILVER.	LEAD.	MERCURY. (Mercurous salts).
HCl hydrochloric acid.	(1) White curdy ppte. of AgCl insoluble in water and HNO_3 ; readily <i>soluble</i> in NH_4HO .	(1) White ppte. of PbCl_2 , soluble in <i>hot</i> water, crystallises out on cooling; <i>unchanged in</i> <i>appearance</i> on addition of NH_4HO .	(1) White ppte. of Hg_2Cl_2 ; insoluble in cold HNO_3 , and <i>black-</i> <i>ened</i> by NH_4HO .
H_2S sulphuretted hydrogen.	(2) Black ppte. of Ag_2S , soluble in boil- ing HNO_3 .	(2) Black ppte. of PbS , soluble in boiling HNO_3 .	(2) Black ppte. of Hg_2S ; not dissolved in boiling HNO_3 .
$(\text{NH}_4)_2\text{S}$ ammonium sulphide.	(3) Same as with H_2S .	(3) Same as with H_2S .	(3) Same as with H_2S .
H_2SO_4 sulphuric acid.		(4) White ppte. of PbSO_4 , soluble in boiling HCl , and the solution, on cooling, deposits needle- shaped crystals of PbCl_2 .	

LEAD, MERCURY, COPPER, IRON, CALCIUM.

MERCURY. (Mercuric salts).	COPPER.	IRON. (Ferrous salts).	IRON. (Ferric salts).	CALCIUM.
(1) When H_2S is added gradually—first white, changing to yellow, then brown, finally black ppte. of HgS , insoluble in HCl , also in HNO_3 ; soluble in aqua regia.	(1) Black ppte. of CuS , soluble in HNO_3 .	No ppte. in neutral or acid solutions.	(1) In acid solutions a white ppte. of sulphur and the ferric is reduced to ferrous salt, $Fe_2Cl_6 + H_2S = 2FeCl_3 + 2HCl + S$.	
(2) Same as with H_2S .	(2) Same as with H_2S .	(2) Black ppte. of FeS , insoluble in alkalies, soluble in HCl .	(2) Black ppte. of FeS , mixed with sulphur, thus:— $Fe_2Cl_6 + S(NH_4)_2S = 2FeS + 6NH_4Cl + S$.	
				(1) White ppte. of $CaSO_4$ from strong solutions of calcium salts.

	SILVER.	LEAD.	MERCURY. (Mercurous salts).
NaHO sodium hydrate or KHO potassium hydrate.	(4) Light brown ppt. of Ag_2O , insoluble in excess of NaHO , soluble in excess of NH_4HO .	(5) White ppt. of $\text{Pb}(\text{HO})_2$, soluble in excess of NaHO .	(4) Black ppt. of Hg_2O , insoluble in excess of NaHO or NH_4HO .
NH_4HO ammonium hydrate.	(5) Same as with NaHO , soluble in excess of NH_4HO .	(6) White ppt. of basic salt, insoluble in excess of NH_4HO .	(5) Black ppt. of basic salt, insoluble in excess of NH_4HO .
K_2CrO_4 potassium chromate.	(6) Dark red ppt. of Ag_2CrO_4 , soluble in hot HNO_3 ; on cooling deposits needle-shaped crystals.	(7) Bright yellow ppt. of PbCrO_4 , readily soluble in NaHO ; with difficulty in HNO_3 .	
KI potassium iodide.		(8) Bright yellow ppt. of PbI_2 , soluble in boiling water, from which it separates on cooling in beautiful golden-yellow scales.	(6) Dark green ppt. of Hg_2I_2 .
SnCl_2 stannous chloride.			(7) Grey ppt. of Hg . On pouring off the fluid and boiling with HCl , the Hg is obtained in little globules.

MERCURY. (Mercuric salts).	COPPER.	IRON. (Ferrous salts).	IRON. (Ferric salts).	CALCIUM.
(3) Yellow ppt. of HgO , insoluble in excess of NaHO .	(3) Pale blue ppt. of $\text{Cu}(\text{HO})_2$, insoluble in excess of NaHO . If NaHO be added in excess and the mixture boiled, the ppt. becomes black.	(3) White ppt. of $\text{Fe}(\text{HO})_2$, which quickly becomes dirty green, and then reddish brown $\text{Fe}_2(\text{HO})_6$.	(3) Reddish brown ppt. of $\text{Fe}_2(\text{HO})_6$, insoluble in excess of the reagent.	
(4) White ppt. of $\text{HgCl}(\text{NH}_4)$ "white precipitate," in solutions of HgCl_2 .	(4) Dark blue solution when NH_4HO is added in excess. Greenish blue ppt. when NH_4HO is added in small quantities.	(4) Same as with NaHO .	(4) Same as with NaHO .	
(5) Bright red ppt. of HgI_2 , soluble in excess either of KI or HgCl_2 .				
(6) SnCl_2 , when added in small quantities to HgCl_2 , gives a white ppt. of Hg_2Cl_2 , which when boiled with excess of the reagent is reduced to grey metallic Hg . Pour off the fluid and boil with HCl ; little globules of Hg are obtained.				

	SILVER.	LEAD.	MERCURY. (Mercurous salts.)
Metallic deposits.			(8) If a strip of bright immersed in a solution metallic mercury is becomes bright on appears on heating
K₄Fe(CN)₆. potassium ferro- cyanide.			
K₃Fe(CN)₆. potassium ferricyanide.			
KCNS potassium sulpho- cyanide.			
(NH₄)₂CO₃ ammonium carbonate.			

MERCURY. (Mercuric salts).	COPPER.	IRON. (Ferrous salts).	IRON. (Ferric salts.)	CALCIUM.
metallic copper be of a salt of mercury, deposited, which rubbing, and dis-	(5) Metallic iron (<i>e.g.</i> a knife blade) or a piece of Sn when immersed in a copper solution, precipitates red metallic copper, especially in the presence of a little free acid.			
	(6) Brown ppt. of $\text{Cu}_2\text{Fe}(\text{ON})_2$, insoluble in dilute acids, but decomposed by KHO .	(5) White ppt., which rapidly becomes blue by oxidation to Prussian blue, $\text{Fe}_4(\text{ON})_6$, insoluble in acids, but decomposed by alkalis.	(5) Prussian blue ppt., insoluble in HCl , soluble in $\text{C}_6\text{H}_8\text{O}_4$, and decomposed by NaHO with separation of Fe_2HO_3 .	
		(6) Dark blue ppt.; (Turnbull's blue), insoluble in acids, decomposed by alkalis.	(6) Changes the colour of the solution to a reddish brown, but no ppt.	
		(7) Produces no coloration.	(7) Blood-red coloration even in dilute solutions.	
				(2) White ppt. of CaCO_3 .

	SILVER.	LEAD.	MERCURY. (Mercurous salts).
C.$(\text{NH}_4)_2\text{O}_4$ ammonium oxalate.			
Blowpipe reactions with NaHCO_3 and KCN on charcoal in re- ducing flame.	(7) Bead, metallic, bright, malleable, so- luble in HNO_3 . On addition of HCl to so- lution, white curdy ppte. of AgCl .	(9) Bead, metallic, malleable, soluble in HNO_3 . On addition of H_2SO_4 to solution, heavy white ppte. of PbSO_4 ; yellow incrusta- tion of PbO on charcoal. Bead makes a mark on paper.	
Flame ! Tests.		(10) In previous test (9), the yellow incrusta- tion of PbO on charcoal disappears on heating in outer blowpipe flame, giving a blue colour to the flame.	
Heated in small dry bulb tube with NaHCO_3.			(9) Grey deposit of tube, which when metallic globules.
Heated in a borax bead be- fore the blow- pipe flame.			

MERCURY. (mercuric salts).	COPPER.	IRON. (Ferrous salts).	IRON. (Ferric salts).	CALCIUM.
				(3) White ppte. of calcium oxalate (C_2CaO_4) even in dilute solutions; soluble in HCl or HNO_3 ; insoluble in acetic or oxalic acid.
	(7) Red metallic particles. — Dissolve particles in HNO_3 and add NH_4HO in excess, yields a deep blue solution.	(8) Black magnetic particles of iron, soluble in HNO_3 and HCl to yellow solution. Add $\text{K}_4\text{Fe}(\text{CN})_6$. — yields a deep blue precipitate.		
	(8) Green coloration when compounds of copper are heated in the Bunsen flame. Copper chloride blue.			(4) Orange red coloration when compounds of calcium are heated in Bunsen flame, especially on moistening the substance with HCl .
cury on sides of bed, appears as				
	(9) Green bead whilst hot; blue on cooling. Becomes colourless in reducing flame.	(9) Yellowish red bead in oxidising flame; bottle green in reducing flame.		

26. REACTIONS OF THE ACIDS.

Carbonic acid, nitric acid, sulphuric acid, hydrochloric acid.

1. CARBONIC ACID, H_2CO_3 .

(a) Carbonates (e.g. Na_2CO_3).

All carbonates are insoluble in water except the carbonates of potassium, sodium and ammonium.

All carbonates effervesce on addition of an acid, owing to the escape of carbon dioxide (CO_2). To a small quantity of a carbonate in a test tube add HCl, and pass the escaping gas, by means of a cork and bent tube (or by simply pouring), into another test tube containing lime water ($Ca(OH)_2$ solution) or baryta water ($Ba(OH)_2$ solution). The solution becomes milky owing to formation of $CaCO_3$ or $BaCO_3$.

(b) Bicarbonates (e.g. $NaHCO_3$).

Bicarbonates are decomposed on heating, and become converted into normal carbonates. Dissolve some $NaHCO_3$ in boiling water; the solution effervesces, CO_2 is given off, and Na_2CO_3 remains.

**2. NITRIC ACID, HNO_3 .**

All nitrates are soluble in water.

(1) To a small quantity of the solution of a nitrate add an equal bulk of strong H_2SO_4 . When the mixture is quite cold pour gently a freshly prepared solution (in cold water) of ferrous sulphate down the side of the test tube. A black ring will form where the two liquids meet. On warming, the ring disappears.

(2) When nitrates are heated with strong H_2SO_4 and copper turnings, the copper is dissolved, forming a green solution of copper nitrate, and the tube is filled with deep red vapours of oxides of nitrogen.

(3) On the addition of a few drops of indigo solution, the blue colour of the indigo disappears on boiling.

3. SULPHURIC ACID, H_2SO_4 .

All sulphates are soluble in water except barium, strontium, calcium, and lead sulphates.

(1.) $BaCl_2$ added to solution of a sulphate produces a white precipitate of $BaSO_4$, insoluble in HCl.

(2.) When a sulphate is fused on charcoal with Na_2CO_3 in the reducing blowpipe flame, a sulphide is formed thus—



If a little HCl be dropped on the fused mass, H_2S is given off which may be recognised by its odour; or, if the mass is placed on a silver coin, and moistened with water, a black stain of Ag_2S is produced.

4. HYDROCHLORIC ACID, HCl .

All chlorides are soluble in water except AgCl , and Hg_2Cl_2 . (PbCl_2 in hot water only.)

(1.) AgNO_3 added to a solution of a chloride produces a white curdy precipitate of AgCl , insoluble in HNO_3 , soluble in NH_4HO .

(2.) Powdered chlorides heated with H_2SO_4 and MnO_2 , yield chlorine gas, which bleaches moistened litmus paper.

27. INTRODUCTORY REMARKS

On the analysis of an unknown substance.

The substance supplied to you will contain not more than one metal and one acid from the following list:—silver, lead, mercury, copper, iron, calcium, carbonic acid, nitric acid, sulphuric acid, hydrochloric acid.

When analysing a substance so as to determine the elements of which it is composed, two methods are adopted, viz.: the *dry* method and the *wet* method.

By the *dry* method various tests are applied to the substance in its dry state, and by these tests we are often enabled to form a very correct idea of its composition.

By the *wet* method the substance is first dissolved in water or acids, and then, by a systematic course of analysis, we are able not only to confirm the dry tests, but to discover with absolute certainty the metal or acid present.

The first step in the analysis of the solution containing dissolved salts of the metals is to divide the metals present into their respective groups by means of the group-reagents.

This division into groups and ultimate complete separation of the metals are based on two facts:—

(1) Some substances are soluble in water, acids or alkalis, while others are insoluble, and may therefore be separated from the soluble substances by filtration.

(2) *All* substances in solution (except sodium salts) unite with *some* reagent to form an insoluble compound.

The metals are divided into five groups, of which the following is a list of those commonly occurring, with their group reagents:—

Group 1. Group reagent, HCl.

Metals precipitated as chlorides on addition of HCl, their chlorides being insoluble in water; they are

Silver, Lead, Mercury (as mercurous salt).

Group 2. Group reagent, H_2S in presence of HCl.

Metals precipitated as sulphides in HCl solutions on addition of H_2S ; they are

Mercury (as mercuric salt), Lead, Bismuth, Copper, Cadmium, Arsenic, Antimony, Tin.

Group 3. Group reagent, $(NH_4)_2S$, in presence of NH_4Cl and NH_4HO .

Metals precipitated as sulphides (or hydrates) in alkaline solutions containing NH_4Cl ; they are

Iron, Nickel, Cobalt, Manganese, Zinc (as sulphides), Aluminium, Chromium (as hydrates).

Group 4. Group reagent, $(NH_4)_2CO_3$, in presence of NH_4Cl and NH_4HO .

Metals whose carbonates are precipitated (after the removal of the metals of the preceding groups) by NH_4Cl , NH_4HO , and $(NH_4)_2CO_3$; they are

Barium, Strontium, Calcium.

Group 5. Metals which are not precipitated by any of the above reagents; they are:—

Magnesium, Potassium, Sodium, Ammonium.

We will now suppose that we have a solution containing dissolved salts of one or two metals belonging to each of the groups.

On the addition of the First Group reagent (HCl), to the clear solution, those metals present which belong to the First Group form insoluble compounds with HCl (viz., chlorides) and fall as a precipitate. The precipitate is collected on a filter, and the soluble substances pass through, to be precipitated in the next group by *the next group-reagent*, and so on through all the groups.

After the separation into groups, they are further separated by the addition of other reagents, and finally distinguished by special tests peculiar to each metal. With this preliminary explanation the following remarks showing how the Tables A, B and C, are used for the detection of *one* metal, and *one* acid, will be more readily understood.

First, we add HCl (1st Group reagent) to a portion of the solution, and if either silver, lead, or mercury (as mercurous salt) be present, it will be converted into a chloride which is insoluble, and falls as a white precipitate which may be collected on a filter. Then, by means of the action of ammonium hydrate on the white precipitate, we are able to decide whether the metal present is silver, lead, or mercury (mercurous salt). See Table A or B.

These metals (Ag, Pb, Hg₂) form the *First Group*. They are the only metals which form with HCl insoluble chlorides.

If neither of the First Group metals were present, but one of the remaining metals of the series—mercury (as mercuric salt), copper, iron, or calcium, we should have no precipitate after adding HCl, because the chloride of mercury (HgCl₂), copper, iron, or calcium, which is formed is soluble in water. Now warm, and add H₂S (HCl + H₂S = 2nd Group reagents) to the same solution; if mercury or copper be present, we get a black precipitate of mercuric sulphide or copper sulphide, because each of those sulphides is insoluble in HCl solutions.

For distinction between mercury and copper. See Table A or B.

Mercury (mercuric salts) and copper are members of the *Second Group*, this group being composed of metals which are precipitated as sulphides in HCl solutions.

If no mercury or copper be present, but either iron or calcium, no precipitate appears on addition of H₂S, because the sulphide of iron or calcium which is formed is soluble in acid solutions. To another portion of the solution we add NH₄Cl, NH₄HO, and (NH₄)₂S, and if iron be present, a black precipitate of ferrous sulphide is formed which (soluble in Group 2) is now insoluble, because the solution is made alkaline by addition of ammonium hydrate, and ferrous sulphide is insoluble in alkaline solutions.

Iron is a member of the *Third Group*, this group being composed of metals which are precipitated as sulphides (or hydrates) in alkaline solutions only.

If iron be not present, but calcium present, there has been hitherto no precipitate on addition of the above reagents, because the calcium chloride and sulphide (or sulphhydrate) formed (Groups 1, 2, and 8) are soluble in alkaline as well as in acid solutions as before mentioned.

To another portion of the solution we add NH_4Cl , NH_4HO , and $(\text{NH}_4)_2\text{CO}_3$; a white precipitate of calcium carbonate falls, insoluble in water. Carbonates are decomposed by acids, hence NH_4HO is added first in order to neutralise any acid which might be present.

Calcium is a member of the Fourth Group, this Group being composed of metals (Ba, Sr, Ca) which remain after removing Groups 1, 2, 8, and whose carbonates are insoluble in water.

28. DIRECTIONS FOR THE ANALYSIS.

The student should proceed in the following order:—

- (1) The preliminary examination, including the dry tests for the metal and acid.
- (2) The preparation of a solution of the substance.
- (8) The wet analysis for the metal.
- (4) The special tests for the acid.

29.—PRELIMINARY EXAMINATION.

Notes should be made at once of the result of each step in the analysis (see p. 60).

- (1) Notice the colour and character of the substance supplied; write its description in your notes, also your inference.

It should be remembered that these colours, &c., only give an approximate idea of the composition of the substance.

The substance is—

<i>Colourless</i> , or <i>white</i> (crystals) ...	Probably Hg , Pb , Ag , or Ca salts. If deliquescent (wet). CaCl_2 , $\text{Ca}(\text{NO}_3)_2$, or $\text{Hg}(\text{NO}_3)_2$.
<i>Black</i> (powder)	Charcoal, CuO , Ag_2O .
<i>Yellow</i> (powder)	Sulphur, PbO , Fe_2Cl_6 (moist lumps).
<i>Red</i> (powder)	HgO , Pb_3O_4 , Fe_2O_3 , Cu_2O .
<i>Green</i> (crystals)	CuCl_2 , FeSO_4 .
<i>Blue</i> (crystals)	CuSO_4 , $\text{Cu}(\text{NO}_3)_2$.

(2) Perform the experiments as given in the first column of the following Table; note what occurs, also your inference as to the probable nature of the substance.

30. PRELIMINARY TESTS FOR THE METAL.

<i>Experiment.</i>	<i>Observation.</i>	<i>Inference.</i>
1. Heat a little of the substance in a dry test tube.	(a) Fumes are given off which condense as a sublimate on sides of tube.	If yellow (or brown drops) indicates <i>Sulphur</i> . If white, <i>Mercury</i> salts. If metallic appearance, <i>Mercury</i> .
	(b) A gas is given off.	Oxygen (rekindles a glowing chip), <i>Nitrates</i> . Red fumes, <i>Nitrates</i> . Carbon dioxide (turns lime water milky), <i>Carbonates</i> .
2. (If white sublimate in 1) Heat a small portion of the substance in a small bulb tube with NaHCO_3 .	Metallic mirror on sides of tube forming globules when rubbed.	<i>Mercury</i> .
3. Heat a little of the substance on charcoal with $\text{NaHCO}_3 + \text{KCN}$ in inner blowpipe flame.	(a) Malleable metallic bead with no incrustation on the charcoal.	<i>Silver</i> (25, Ag, 7.)* <i>Lead</i> (25, Pb, 9.) <i>Copper</i> (25, Cu, 7.) <i>Iron</i> (25, Fe, 8.)
	(b) Malleable metallic bead which makes a mark on paper; also yellow incrustation on charcoal.	
	(c) Red beads or scales.	
	(d) Black magnetic powder.	
4. Moisten a fragment of the substance with HCl on the loop of a clean platinum wire,† and hold in the outside edge of lower part of Bunsen flame.	Flame is coloured Green	<i>Copper</i> .
	„ Orange Red	<i>Calcium</i> .
	„ Blue	<i>Copper Chloride</i> , or <i>Lead</i> .

* 25. Ag, 7 means Paragraph 25, silver column, test number 7. Confirm by these tests.

† The wire is cleaned by dipping it in HCl and heating in the Bunsen flame until it no longer colours the flame.

31. PRELIMINARY TESTS FOR THE ACID.

<i>Experiment.</i>	<i>Observation.</i>	<i>Inference.</i>
(1) Add HCl to a little of the substance in a test tube.	Effervescence, and the gas which comes off turns lime water milky.	<i>Carbonic acid.</i>
(2) Treat a small portion of the substance in a test tube with strong H_2SO_4 , and warm gently.	Light red fumes, which condense as yellow drops on sides of tube.	<i>Nitric acid.</i>
	White fumes, acid to test paper; confirm for HCl by adding MnO_2 ; Cl given off, which bleaches litmus paper.	<i>Hydrochloric acid.</i>
	No action.	<i>Sulphuric acid.</i>

32. Preparation of a solution of the substance.

Powder, if necessary, in a clean mortar. The substance is soluble in water, or dilute acids. With separate small portions in a test tube, test its solubility in the following order :—

- (1) In water.
- (2) In hydrochloric acid.
- (3) In nitric acid.
- (4) In a mixture of hydrochloric and nitric acids.

In each case the substance should be boiled for some time, unless it dissolves at once. When the proper solvent has been found, dissolve a larger quantity of the substance and proceed to test for the metal by Table A. Use small portions of your solution *at a time.*

TABLE A.
33. TO DETECT THE METAL.
 Table for the examination of a solution containing one of the following Metals :—Silver, Lead, Mercury, Copper, Iron or Calcium.

To a portion of the solution add HCl. If no ppte., pass on to next column.	Warm the same portion and add to the acid solution H₂S either gas or strong solution. If no ppte., or only a white cloud (sulphur), pass on to next column.	To a fresh portion add NH₄Cl NH₄HO (NH₄)₂CO₃	To a fresh portion add NH₄Cl NH₄HO (NH₄)₂CO₃
Group I. A white ppte. indicates one of the following :— AgCl; Hg₂Cl₂ PbCl₂ Filter. To the white residue add NH₄HO.	Group II. A black ppte. indicates one of the following :— HgS CuS To a portion of the original solution add NH₄HO in excess.	Group III. A black ppte. indicates— FeS Test original solution for Ferrous or Ferric Salt by K₄FeO₆N₃.	Group IV. A white ppte. indicates— CaCO₃ Test dilute original solution with ammonium oxalate.
The ppte. is dissolved indicates SILVER. Confirm by Ag⁺ 6, 7."	The ppte. is unchanged in appearance, indicates LEAD. as mercurous salt. Confirm by Hg₂ 6, 8, 9.	White ppte. indicates MERCURY as mercuric salt. Confirm by Hg. 6, 7, 9.	Dark blue ppte. quickly becoming dark blue indicates IRON as ferric salt.
The ppte. is indicates SILVER. Confirm by Ag⁺ 6, 7."	The ppte. is unchanged in appearance, indicates LEAD. as mercurous salt. Confirm by Hg₂ 6, 8, 9.	Light blue ppte. quickly becoming dark blue indicates IRON as ferrous salt.	Dark blue ppte. indicates IRON as ferric salt.
		indicates CALCIUM. Confirm by flame test.	

TABLE C.

35. TO DETECT THE ACID,

Not more than one of the following being present :—Carbonic acid, nitric acid, sulphuric acid, hydrochloric acid.

Note.—The substance given is soluble in water or dilute acids ; the following insoluble substances are therefore absent,—silver chloride, mercurous chloride, lead sulphate, calcium sulphate (lead chloride is soluble in hot water, but is deposited again on cooling). Hence if silver or mercury (as mercurous salt) be found, do not expect to find hydrochloric acid ; or if lead or calcium be found, do not expect to find sulphuric acid.

Use separate small portions for the following tests :—

(1) *Carbonic acid*, if present, will have been already found by the preliminary tests.

(2) *Nitric acid*. The substance is soluble in water. Apply the special tests for nitric acid (page 46) to the original solution.

(3) *Sulphuric acid*. To a small portion of the original solution add HCl and BaCl₂ ; or, if Ag, Hg₂, or Pb be present, add instead, HNO₃ and Ba(NO₃)₂ ; a white precipitate indicates sulphuric acid.

(4) *Hydrochloric acid*. To a small portion of the original solution* add silver nitrate (AgNO₃) ; a white curdy precipitate, insoluble in HNO₃, soluble in NH₄HO, indicates hydrochloric acid.

36. QUESTIONS ON THE REACTIONS OF THE METALS.

I.

(1). What is the result of adding HCl to solutions of silver nitrate, lead nitrate, and mercurous nitrate ; and how would you distinguish between these solutions supposing they were not labelled ?

(2) By means of HCl we are able to divide a mixture of the eight solutions mentioned in the table into two groups ; explain why this can be done.

(3) Silver chloride is dissolved in NH₄HO and nitric acid is afterwards added in excess. What is the result ; and why is the acid added in excess ?

(4) Write the equations showing the action of HCl on the nitrates in question (1).

II.

(1) Having mixed together solutions of lead, mercury (mercurous salt) and copper, how would you separate the metals present ?

* Care should be taken that the original substance was dissolved without addition of HCl.

(2) On adding water to a white powder and boiling, the solution becomes quite clear. After a few minutes, on looking at the solution a crystalline residue is seen. Explain this.

(3) Mention two methods of separating lead from silver.

(4) Write out the formulæ for the chlorides of each of the metals in the Table.

III.

(1) Write out a Table showing the action of H_2S on solutions of silver, lead, mercury, copper, iron, and calcium.

(2) How may sulphide of mercury be separated from the sulphides of silver, lead, and copper?

(3) How may ferric be distinguished from ferrous salts by means of H_2S in acid solutions?

(4) Write equations showing the action of H_2S on solutions of the nitrates of silver, lead, and mercury (mercurous and mercuric salts).

IV.

(1) Why is there no precipitate when H_2S is added to acid solutions of ferrous salts? (See sulphuretted hydrogen, page 49.)

(2) On the addition of H_2S to a solution, the precipitate formed was at first white, then yellow, brown, and at length black; what metal is present, and how may the black precipitate be dissolved?

(3) What is the result of adding ammonium sulphide to solutions of lead, copper, and calcium salts?

V.

(1) Write out the results of adding $NaHO$ to solutions of silver, lead, mercury, copper, iron, and calcium salts. How would you distinguish between ferrous and ferric salts, also between mercurous and mercuric salts by means of potassium hydrate?

(2) How do you explain the fact that pale blue copper hydrate, on boiling with excess of $NaHO$, becomes black?

(3) A solution contains dissolved salts of lead and copper; how would you separate them?

(4) Write equations showing the results of adding $NaHO$ to solutions of lead nitrate, ferrous sulphate, ferric chloride, mercurous nitrate, and mercuric chloride.

VI.

- (1) What is "white precipitate," and how is it obtained?
- (2) What is the action of NH_4HO on a solution of a copper salt?
- (3) How would you distinguish between silver and lead by means of potassium chromate; also between mercurous and mercuric salts by means of stannous chloride?
- (4) Write equations showing the action of K_2CrO_4 on silver nitrate and lead nitrate; also of SnCl_2 on mercurous and mercuric salts.

VII.

- (1) How would you distinguish between lead, mercurous, and mercuric salts by means of potassium iodide?
- (2) What would be the effect of adding sulphuric acid to solutions of lead nitrate, copper chloride, and calcium chloride? Give equations.
- (3) A strip of bright copper is placed in a solution of mercuric chloride; what is the result? Give an equation.
- (4) How would you test for the presence of metallic copper in a solution of copper sulphate? Give an equation.

VIII.

- (1) Potassium ferrocyanide when added to three separate solutions gave a brown, a white changing rapidly to deep blue, and a deep blue precipitate respectively. Say what metal is indicated in each case; also how would you confirm your inference?
- (2) Write, side by side, a list of those tests which enable you to distinguish between ferrous and ferric salts.
- (3) What is the result of adding $(\text{NH}_4)_2\text{CO}_3$ to neutral solutions of silver, lead, mercury, copper, iron, and calcium salts; and what precaution is necessary before concluding that a white precipitate with $(\text{NH}_4)_2\text{CO}_3$ indicates presence of the Calcium Group.
- (4) You have given you a *dilute* solution of calcium chloride? What test is used for detecting the calcium? Give an equation.

IX.

- (1) Write out the blowpipe tests for silver, lead, copper, and iron.
- (2) How would you distinguish between a particle of silver and a particle of lead; also between particles of copper and iron?

(3) What are the flame tests for copper and calcium; also the borax bead tests for copper and iron?

(4) Write, side by side, a list of the tests which enable us to distinguish between mercurous and mercuric salts.

X.

(1) Name the metals present in the following solutions:—

(a) On addition of HCl , a white precipitate soluble in NH_4HO .

(b) On addition of NH_4HO in excess, solution becomes of a deep blue colour.

(c) On addition of NaHO , a pale blue precipitate, which on boiling with excess of NaHO becomes black.

(d) On addition of H_2S , a black precipitate, which is not dissolved after boiling with HNO_3 .

(e) On addition of HCl , a white precipitate, which is soluble in hot water, and crystallises out on cooling.

(f) On addition of HCl and H_2S , no precipitate, but on the further addition of ammonia a black precipitate. The black precipitate dissolves on addition of HCl with evolution of H_2S gas; and a few drops of potassium ferrocyanide added to the clear solution gives a white or light blue precipitate which rapidly becomes darker.

(g) On addition of mercuric chloride, a bright red precipitate, which disappears on addition of more mercuric chloride.

(h) On addition of mercuric chloride, a white precipitate, which, when boiled with excess of mercuric chloride, becomes grey.

XI.

Name the substances which behave as follows:—

(a) The substance is white powder, it does not dissolve in water; on adding HCl a rapid effervescence occurs, and the gas which comes off turns lime water milky; flame is coloured orange red; the metals of the first three groups are absent, and on addition of NH_4Cl , NH_4HO , and $(\text{NH}_4)_2\text{CO}_3$ to a solution of the substance dissolved in HCl , a white precipitate falls.

(b) Substance dissolves in water, gives off red fumes on heating, addition of sulphuric acid to solution produces a white precipitate; addition of H_2S gives a black precipitate. On adding a little strong H_2SO_4 , and FeSO_4 solution, a black ring appears where the two solutions meet.

(c) Substance is blue, dissolves in water, solution gives a black precipitate with H_2S , and a white precipitate after adding HCl and BaCl_2 , NH_4HO added in excess gives deep blue solution, flame reaction green.

(d) Substance is white, soluble in water, gives no precipitate on addition of HCl to solution; on heating in a bulb tube with NaHCO_3 a metallic mirror appears, which may be rubbed into little globules; solution gives a white curdy precipitate with AgNO_3 , soluble in NH_4HO .

(e) Substance is a red heavy powder, gives off oxygen when heated in a tube, and becomes yellow; on heating before the blow-pipe on charcoal, a yellow incrustation forms on the charcoal and a metallic bead appears, the bead is soluble in HNO_3 , and the solution gives a white precipitate with H_2SO_4 .

XII.

(1) Name the substances which behave as follows:—

(a) The substance is green and crystalline, readily dissolves in water, on adding the solution to a mixture of nitric and strong sulphuric acids, a black ring appears where the liquids meet.

(b) The substance is a black powder, colours the flame green, dissolves readily in H_2SO_4 to form a blue solution.

(2) What are bi-carbonates, and what is the result of dissolving sodium bi-carbonate in boiling water?

(3) Write out the tests for nitrates.

(4) How are sulphates detected, (i) in the dry way (ii) in the wet way?

XIII.

(1) How would you test for the presence of a chloride?

(2) Write out a list of the commonly occurring metals, arrange them in groups and give the group reagents.

(3) Give the reason why the metals of each group are placed in that group.

(4) You have a powdered substance to analyse. How would you proceed to prepare a solution of the substance?

XIV.

(1) Write out and learn the preliminary tests for the metals.

(2) Write out and learn the preliminary tests for the acids.

XV.

(1) Write out and learn Table A.

(2) Write out and learn Tables B and C.

39. How to write down the result of your Analysis.*An actual example copied from a student's note book.***Description:—**Blue crystalline substance—probably Copper Sulphate, or Nitrate.**Solubility:—**Soluble in water.**PRELIMINARY TESTS FOR METAL.**

<i>Experiment.</i>	<i>Observation.</i>	<i>Inference.</i>
Heated powdered substance in dry test tube.	Liquefied; water in tube acid to test paper. When all water driven off, red fumes appeared	Water of crystallisation. Nitrate.
Heated on charcoal with NaHCO_3 and KCN in reducing flame.	Red scales.	Copper.
Heated on platinum wire in outer edge of Bunsen flame.	Flame is coloured green.	Copper.

PRELIMINARY TESTS FOR ACID.

Heated in test tube with strong H_2SO_4 .	Red fumes.	Nitric Acid.
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TO FIND THE METAL.

To a portion of the solution added HCl.	No ppte.	No 1st Group.
To the same portion added H_2S .	Black ppte.	Copper or Mercury.
To original solution added NH_4HO .	Deep blue solution.	Copper.
Confirmed with knife blade in solution.	Red deposit.	Copper.

TO FIND THE ACID.

Added to solution strong H_2SO_4 , and when cold, FeSO_4 .	Black ring.	Nitric Acid.
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Result..... { **Copper.**
 { **Nitric Acid.**

PART II.

NOTES AND QUESTIONS
ON
THEORETICAL CHEMISTRY.

PART II.

THEORETICAL CHEMISTRY.

NOTES.

TABLE OF THE ELEMENTS.

NON-METALS.

NAME.	Symbol.	Atomic Weight.	NAME.	Symbol.	Atomic Weight.
Oxygen	O	16	Carbon.....	C	12
Hydrogen	H	1	Sulphur	S	32
Chlorine	Cl	35.5	Phosphorus....	P	31
Bromine	Br	80	Boron	B	11
Iodine	I	127	Silicon	Si	28
Fluorine	F	19	Selenium.....	Se	79
Nitrogen	N	14	Tellurium	Te	128

METALS.

NAME.	Symbol.	Atomic Weight.	NAME.	Symbol.	Atomic Weight.
Antimony	Sb	122	Magnesium....	Mg	24
Arsenic	As	75	Potassium	K	39
Tin	Sn	118	Sodium	Na	23
Lead	Pb	207	Iron	Fe	56
Silver	Ag	108	Aluminium	Al	27.3
Mercury	Hg	200	Chromium	Cr	52.5
Bismuth	Bi	208	Manganese	Mn	55
Copper.....	Cu	63.5	Zinc	Zn	65
Cadmium	Cd	112	Nickel	Ni	58.8
Barium	Ba	137	Cobalt	Co	58.8
Strontium	Sr	87.5	Gold.....	Au	196.7
Calcium	Ca	40	Platinum	Pt	197.4

A **simple substance, or element**, is one which cannot be split up by any known process into two or more different substances, *e.g.*, Hydrogen, Oxygen.

A **compound substance** is one which is produced by the union (in certain exact proportions) of two or more simple substances, and therefore can be split up into two or more different substances.

An **atom** is the smallest possible particle of an *element*.

A **molecule** is composed of atoms, and is the smallest possible particle of either an element or a compound *which can exist alone*.

The **atomic weight** of an element is the weight of an atom of that element compared with the atom of hydrogen.

The **molecular weight** of a compound is equal to the sum of the atomic weights of its constituent elements.

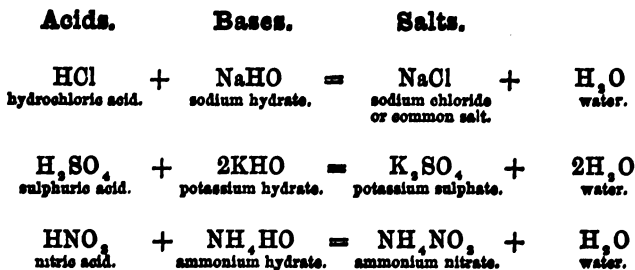
An **acid** is a compound containing one or more atoms of hydrogen, which may be displaced by a metal when the metal is presented to it in the form of a hydrate. Acids possess a sour taste, and turn blue litmus solution red; they cause effervescence when added to carbonates, and they neutralise alkalies.

A **base** is a substance which will neutralise an acid.

A **salt** is the compound formed by the union of an acid and a base.

Alkalies are bases which are soluble in water. They readily neutralise acids and restore the blue colour to reddened litmus solution. The principal alkalies are sodium hydrate, NaHO ; *potassium* hydrate, KHO ; ammonium hydrate, NH_4HO .

Examples of the formation of salts from acids and bases :—



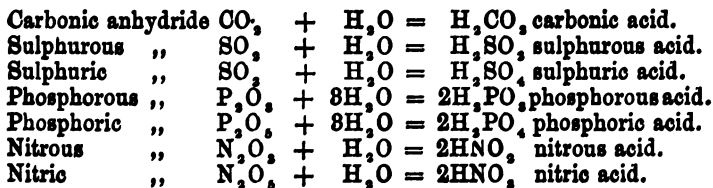
NOTE.—Hydrochloric acid HCl.. forms chlorides e.g. NaCl sodium chloride.
 Sulphuric acid H₂SO₄.. forms sulphates e.g. Na₂SO₄ sodium sulphate.
 Sulphurous acid H₂SO₃.. forms sulphites e.g. Na₂SO₃ sodium sulphite.
 Nitric acid HNO₃..... forms nitrates e.g. NaNO₃ sodium nitrate.
 Nitrous acid HNO₂..... forms nitrites e.g. NaNO₂ sodium nitrite.
 Phosphoric acid H₃PO₄.. forms phosphates .. e.g. Na₃PO₄ trisodium phosphate
 Phosphorous acid H₃PO₃.. forms phosphites .. e.g. Na₂PO₃ trisodium phosphite.

Acids ending in *ic* form salts ending in *ate*.

" " *ous* " " *ite*.

Salts which contain no oxygen end in *ide*.

Anhydrides. The word *anhydride* signifies *free from water*. The name is applied to certain oxides, which when added to water form acids, thus :—



ATOMICITY OF THE ELEMENTS.

Elements are divided into monads, dyads, triads, tetrads, pentads and hexads, according to the number of their *bonds*, that is, according to the number of atoms of hydrogen or chlorine they will replace, or combine with, whether it be one, two, three, four, five or six, thus :—

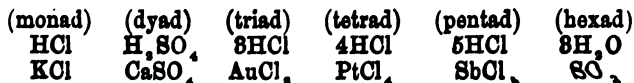


TABLE OF ATOMICITIES OF THE ELEMENTS:

MONADS.	DYADS.	TRIADS.	TETRADES.	PENTADS.	HEXADS.
Hydrogen. Chlorine. Bromine. Iodine. Fluorine. Silver. Potassium. Sodium.	Oxygen. Barium. Strontium. Calcium. Magnesium. Mercury. Zinc. Cadmium. Copper. Lead.	Boron. Gold. Bismuth. Antimony. Arsenic. Nitrogen. Phosphorus.	Carbon. Silicon. Platinum. Tin. Aluminium. Lead.	Bismuth. Antimony. Arsenic. Nitrogen. Phosphorus.	Sulphur. Iron. Chromium. Manganese. Nickel. Cobalt.

It will be noticed that some elements appear as triads and pentads (odd numbers), others as dyads and tetrads (even numbers), but none as dyads and triads (both odd and even).

CHEMICAL COMBINATION IN DEFINITE AND MULTIPLE PROPORTIONS.

(I.) Bodies combine chemically in *definite* proportions. When two elements unite, one atom of the first combines with one, two, or three atoms of the second, as the case may be, but never with any fraction of an atom (which is impossible by the definition of an atom). When hydrogen and oxygen unite to form water, they do so in the proportion of *two* volumes of hydrogen to *one* volume of oxygen, and never in any other proportion. Any excess of either gas would remain behind uncombined.

Again, when mercury combines with oxygen, 200 grams of mercury combine with *exactly* 16 grams of oxygen. If 17 grams of oxygen were present, one gram of oxygen would remain behind uncombined.

The same principle applies in the formation of all chemical compounds.

(II.) *Multiple proportions.* Some substances unite to form more than one compound, thus :—There are two compounds of carbon and oxygen; in one, 12 grams of carbon unite with 16 grams of oxygen, forming carbon monoxide; in the other, 12 grams of carbon unite, not with 15 or 17, but with exactly 32 grams of oxygen, forming carbon dioxide. Hence the oxygen in carbon dioxide is a simple multiple of the oxygen in carbon monoxide.

Again, nitrogen forms five different compounds with oxygen, and the weights of the combined oxygen are proportional to the numbers, 1, 2, 3, 4, 5, and each is a simple multiple of the first, thus :—

28	parts of nitrogen unite with	16	parts of oxygen, forming nitrogen monoxide, N_2O .
28	"	32	" forming nitrogen dioxide, N_2O_2 .
28	"	48	" forming nitrogen trioxide, N_2O_3 .
28	"	64	" forming nitrogen tetroxide, N_2O_4 .
28	"	80	" forming nitrogen pentoxide, N_2O_5 .

Here the numbers 16, 32, 48, 64, 80 divided by 16 give 1, 2, 3, 4, 5.

DALTON'S ATOMIC THEORY.

In seeking to discover the reason why substances combined in these definite or simple multiple proportions, Dalton supposed :—

- (i.) That all matter is composed of extremely small particles which he called *atoms* (so called because it was assumed they could not be in any way cut or divided).
- (ii.) That the atoms of the same element are alike in every respect.
- (iii.) That the atoms of different elements differ in character and properties, and bear the same relation to one another by weight as the combining weights of these elements.
- (iv.) That the formation of chemical compounds merely consists in the union or approximation of the atoms of different elements to one another.

Differences between a Mechanical Mixture and a Chemical Compound.

Mechanical Mixture.

Example :—Iron and Sulphur merely mixed ($\text{Fe} + \text{S}$).

1. The component parts are present in *any* proportion.
2. The properties of a mixture are the *mean* of the properties of its constituents.
3. When the components are mixed together *no heat* is evolved, because no union takes place between them.
4. When examined under a microscope, the component parts are readily distinguished one from another.
5. The component parts may be separated by the aid of water or other solvent.

Chemical Compound.

Example :—Iron and Sulphur mixed together in the proportion of 56 by weight of Fe to 32 of S and then heated, unite to form a new compound, viz., Ferrous Sulphide (FeS).

1. The component parts are present in certain *exact* proportions.
2. The properties of a compound are *entirely different* from the properties of its constituents.
3. When the components combine together *heat* is evolved, due to their union to form a new compound.
4. When examined under a microscope the substance presents a *uniform* appearance throughout.
5. The component parts cannot be separated by the aid of water or any solvent.

PRELIMINARY NOTE ON SIMPLE PROPORTION, OR RULE OF THREE.

Rule :—Multiply the 2nd and 3rd terms of a proportion together, and divide by the 1st term; the result gives us the 4th term, or answer.

Example :—If 18 parts of water contain 16 parts of oxygen, *how many parts of oxygen* will 100 parts of water contain?

1st term.		2nd term.		3rd term.		4th term.
water.		water.		oxygen.		
18	:	100	::	16	:	answer.
				100		
				18)1600(88·9 nearly		
				144		
				160		
				144		
				160		

88·9 parts. Answer.

With this kind of sum always ask yourself two questions :—

(i.) **Of what will the answer consist?** Look at your question and you find the words—*How many parts of oxygen?* Then put the *known* quantity of oxygen in the 3rd term (see above).

(ii.) **Will the answer be greater or less than the number now in the third term?** There will be more oxygen in 100 parts of water than in 18 parts; therefore so arrange your proportion that the answer shall be more.

To find the molecular weight of a substance.

Example :—Find the molecular weight of copper nitrate $\text{Cu}(\text{NO}_3)_2$.

NOTE.—The molecular weight, or the weight of a molecule, of a substance is equal to the total weight of the atoms composing it.

$$\begin{array}{rcl}
 \text{Cu} & = & 63\cdot5 \\
 2\text{N} & = & 14 \times 2 = 28 \\
 6\text{O} & = & 16 \times 6 = 96 \\
 & & \hline
 & & 187\cdot5
 \end{array}$$

To find the percentage composition of a substance.

Example :—Find the percentage composition of KClO_3 (potassium chlorate).

NOTE.—The question requires you to find how many parts of potassium, chlorine, and oxygen, by weight, are contained in 100 parts of KClO_3 .

- (1) Find the molecular weight of the compound.
- (2) Find by proportion the amount of each element in every 100 parts of the compound.

$$\text{Molecular weight of } \text{KClO}_3 = 39 + 35.5 + 48 = 122.5.$$

Weights of K, Cl, and O in 122.5 parts of KClO_3 are known, therefore for the required answers we arrange the following proportions :—

KClO_3 122.5	:	KClO_3 100	::	Potassium. 89	:	x
KClO_3 122.5	:	KClO_3 100	::	Chlorine. 35.5	:	x
KClO_3 122.5	:	KClO_3 100	::	Oxygen. 48	:	x

On working out these proportions we get :—

$$\left. \begin{array}{l} \text{Potassium} \dots 31.84 \\ \text{Chlorine} \dots\dots 28.98 \\ \text{Oxygen} \dots\dots 39.18 \end{array} \right\} = \text{percentage composition of } \text{KClO}_3.$$

$$100.00$$

PROBLEMS INVOLVING QUANTITIES OF SUBSTANCES BY WEIGHT.

Rule 1. Write out the equation.

Rule 2. Put the atomic or molecular weights under each term of the equation. You can then reason from the known to the unknown, and obtain the required answer by a simple proportion.

Examples showing the various forms in which a question may be presented:—

- (i.) How much Na combines with 100 grams of Cl ?
 (ii.) " " Cl " " 100 grams of Na ?
 (iii.) " " NaCl can be formed from 100 grams of Cl ?
 (iv.) " " NaCl " " 100 grams of Na ?
 (v.) " " Na is contained in 100 grams of NaCl ?
 (vi.) " " Cl " " 100 grams of NaCl ?

Solution of (i.)—



(b) The atomic or } $23 + 85.5 = 58.5$
 molecular weights }

(c) The answer will consist of Na (the question is—*How much Na?*) Put Na with its weight in the 3rd term.

(d) The answer will be *greater* than 23 (the number in the 3rd term), because there will be more Na in 100 parts of NaCl than in 58.5 parts—see equation in (b)
 —Hence arrange your proportion so that the answer shall be greater, thus:—

$$\begin{array}{rclclcl}
 \text{Cl} & & \text{Cl} & & \text{Na} & & \\
 85.5 & : & 100 & :: & 23 & : & x \\
 & & & & 100 & & \\
 & & & & 85.5 \overline{)2300} & (64.78 & \\
 & & & & 2180 & & \\
 & & & & \underline{1700} & & \\
 & & & & 1420 & & \\
 & & & & \underline{2300} & & \\
 & & & & 2485 & & \\
 & & & & \underline{8150} & &
 \end{array}$$

Amount of Na required to unite } = 64.78 grams Ans.
 with 100 grams of Cl }

The following are the proportions for the solutions of (ii.), (iii.), (iv.), (v.), (vi.):—

	Na	:	Na	::	Cl	:	x
(ii.)	23	:	100	::	85.5	:	x
	Cl	:	Cl		NaCl		
(iii.)	85.5	:	100	::	58.5	:	x
	Na	:	Na		NaCl		
(iv.)	23	:	100	::	58.5	:	x
	NaCl	:	NaCl		Na		
(v.)	58.5	:	100	::	23	:	x
	NaCl	:	NaCl		Cl		
(vi.)	58.5	:	100	::	85.5	:	x

ON THE LAW OF VOLUMES.

2 atoms of H + 1 atom of O = 1 molecule of H_2O .

The above may be read thus :—

2 volumes of H + 1 volume of O = 2 volumes of H_2O .

From this it will be seen that for atom, we read *volume*, and for molecule, *two volumes*. Now note particularly :—No matter how many atoms of different elements may be required in order to form a compound, when they unite they form *one molecule* of that compound ; or, in other words,—No matter how many volumes of different elements may be required in order to form a compound, the resulting compound always occupies only *two volumes*. These facts have been determined by experiment. Examples :—

1 vol. of H + 1 vol. of Cl = 2 vols. of HCl.

2 vols. of H + 1 vol. of S = 2 vols. of H_2S .

8 vols. of H + 1 vol. of N = 2 vols. of NH_3 .

4 vols. of H + 1 vol. of C = 2 vols. of CH_4 .

2 vols. of N + 1 vol. of O = 2 vols. of N_2O .

To find the weight of a given volume of a simple or compound gas at 0°C . and 760 m.m. pressure.

NOTE.—The weight of 1 litre of hydrogen at the normal temperature and pressure, viz :—At 0°C . and 760 m.m. of mercury = .0896 grams, or 1 crith. Remembering this, we are able to calculate the weight of any gas whether simple or compound.

I. Simple or elementary gases.

Example :—Find the weight of 1 litre of oxygen at N.T.P.*

Now the atomic weight of oxygen is 16 ; that is, oxygen is 16 times as heavy as hydrogen.

But 1 litre of hydrogen weighs .0896 grams.

∴ 1 litre of oxygen weighs $.0896 \times 16 = 1.43$ grams.

In the same way—

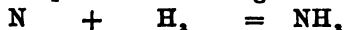
1 litre of nitrogen weighs $.0896 \times 14 = 1.25$ grams.

1 litre of chlorine weighs $.0896 \times 35.5 = 3.18$ grams.

*NOTE. N.T.P.=Normal temperature and pressure, viz. :— 0°C . and 760 m.m. pressure.

II. Compound gases.

Example :—Find the weight of 1 litre of ammonia at N.T.P.



1 volume + 3 volumes = 2 volumes.

14 + 8 = 17 = weight of 2 volumes.

Therefore weight of 1 volume = $\frac{17}{2} = 8.5$ (compared with hydrogen)

But weight of 1 litre of H = .0896 grams.

$\therefore$ weight of 1 litre of NH_3 = $.0896 \times 8.5$ grams.
= .7616 grams.

To find the volume occupied by a given weight of a simple or compound gas at 0°C . and 760 m.m. pressure.

We know that .0896 gram of H occupies 1 litre (at N.T.P.).

(Dividing through-
out by .0896.) $\therefore$ 1 gram ,, $\frac{1}{.0896}$ litre.
= 11.16 litres.

Remembering this fact, we can calculate the volume occupied by any gas when we know its weight compared with hydrogen at the same temperature and pressure, thus :—

Because 1 gram of H occupies 11.16 litres.

$\therefore$ 16 grams of O occupy 11.16 ,,

Also 14 ,, N ,, 11.16 ,,

85.5 ,, Cl ,, 11.16 ,,

82 ,, S vapour ,, 11.16 ,,

For compound gases, since the weight of one volume compared with hydrogen = $\frac{\text{molecular weight}}{2}$, therefore—

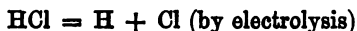
$\frac{17}{2}$ or 8.5 grams of NH_3 occupy 11.16 litres.

$\frac{18}{2}$ or 9 ,, H_2O ,, 11.16 ,,

$\frac{36.5}{2}$ or 18.25 ,, HCl ,, 11.16 ,,

$\frac{64}{2}$ or 32 ,, SO_2 ,, 11.16 ,,

Example :—What volume of chlorine can be obtained from 50 grams of hydrochloric acid at N.T.P. ?



$$86.5 = 1 + 85.5$$

Now 35.5 grams of chlorine occupy 11.16 litres.

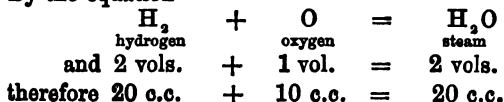
Hence the question becomes—If 11.16 litres of chlorine can be obtained from 86.5 grams of hydrochloric acid, how many litres of chlorine can be obtained from 50 grams of that acid?

grams	:	grams	::	litres	:	x
86.5	:	50	::	11.16	:	x
$\frac{50 \times 11.16}{86.5} = 15.287 \text{ litres of chlorine.}$						

To find the volume resulting from the combination of gaseous bodies.

Example 1.—If a mixture of 20 c.c. of hydrogen and 10 c.c. of oxygen at 140°C . be exploded, what volume of steam will be formed?

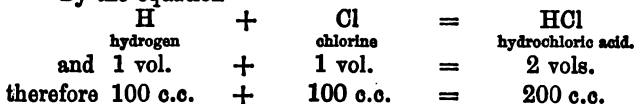
By the equation



Answer. 20 c.c. of steam.

Example 2.—If 100 c.c. of hydrogen are mixed with 150 c.c. of chlorine, and exploded, what volume of hydrochloric acid is produced? Which gas, and how much of it, remains uncombined?

By the equation—

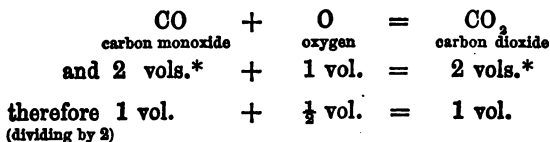


From this it will be seen that 100 c.c. of the chlorine unite with *all* the hydrogen forming 200 c.c. of hydrochloric acid gas; the remaining 50 c.c. of chlorine are left uncombined.

Answer. 200 c.c. of HCl ; 50 c.c. Cl .

Example 8.—One volume of carbon monoxide (carbonic oxide) gas is mixed with two volumes of oxygen, and an electric spark passed through the mixture. What volume of carbon dioxide gas (carbonic anhydride) is formed, and what volume of oxygen left? (Science and Art Department Examination. 1877.)

By the equation—



From this it will be seen that 1 volume of carbon monoxide requires only $\frac{1}{2}$ volume of oxygen to unite with it to form carbon dioxide; the additional $1\frac{1}{2}$ volumes of oxygen remain behind uncombined.

Answer : 1 volume of carbon dioxide ; $1\frac{1}{2}$ volumes of oxygen left.

* It should be remembered that a molecule is always written 2 volumes, whatever number of atoms it contains.

Density and specific gravity of gases.

The **density** of a gas = its weight compared with hydrogen ; and, therefore, for elementary gases, = its atomic weight, and for compound gases, = $\frac{\text{its molecular weight.}}{2}$

Thus, density of oxygen		= 16
„ nitrogen		= 14
„ steam		= $\frac{18}{2} = 9$
„ carbon dioxide		= $\frac{44}{2} = 22$

The **specific gravity** of a gas = its weight compared with air. Air is 14.42 times as heavy as hydrogen, therefore the weight of hydrogen compared with air is = $\frac{1}{14.42} = .0698$.

Hence specific gravity of hydrogen = .0698.

Remembering the specific gravity of hydrogen, it is easy to calculate the specific gravity of any gas, viz.:—By multiplying $\cdot 0698$ by the weight of one volume of the gas compared with hydrogen.

Thus, specific gravity of oxygen, compared with air, is—

$$\cdot 0698 \times 16 = 1\cdot 108$$

$$\text{of chlorine} \dots\dots\dots \cdot 0698 \times 85\cdot 5 = 2\cdot 460$$

$$\text{of hydrochloric acid} \cdot 0698 \times 18\cdot 25 = 1\cdot 264$$

Empirical formulæ express (1) the elements composing a compound, (2) the *proportional* number of atoms of each element present in a molecule of the compound; thus, CH_3O is the empirical formula for acetic acid.

Rational formulæ express the *actual* number of atoms of each element present in a molecule of a compound; thus, $\text{C}_2\text{H}_4\text{O}_2$ is the rational formula for acetic acid.

The empirical formula for a compound is used only when the rational formula is unknown.

To find the empirical formula of a substance from its percentage composition.

Rule :—(a) Divide the percentage amount of each element by its atomic weight.

(b) Divide each of the results in (a) by the lowest result.

(c) Reduce to simplest relation in whole numbers.

Example :—A compound consists of iron and oxygen in the following proportion :—

Iron	70·00
Oxygen	80·00
	100·00

Atomic weight of Fe = 56; atomic weight of O = 16

$$\text{By (a) Iron} = \frac{70}{56} = 1\cdot 25$$

$$\text{Oxygen} = \frac{80}{16} = 1\cdot 875$$

$$\text{By (b) Iron} = \frac{1\cdot 25}{1\cdot 25} = 1$$

$$\text{Oxygen} = \frac{1\cdot 875}{1\cdot 25} = 1\cdot 5 \text{ (or } 1\frac{1}{2}\text{)}$$

By (c) We now find the lowest whole numbers in the ratio of $1 : 1\frac{1}{2}$, as we cannot have fractions of atoms. Multiplying each by 2, we get—



Hence formula = $\text{Fe}_2 \text{O}_3$ (ferric oxide).

SUMMARY OF EQUATIONS.

You are recommended to use these equations in the following way:—After having carefully learnt them, test your knowledge by covering the right hand side with a strip of paper, and complete the equation from memory.

HYDROGEN.

Preparation.

1. K_2 + $2\text{H}_2\text{O}$ = 2KHO + H_2
potassium. water. caustic potash. hydrogen.
2. Na_2 + $2\text{H}_2\text{O}$ = 2NaHO + H_2
sodium. water. caustic soda. hydrogen.
3. Zn + H_2SO_4 = ZnSO_4 + H_2
zinc. sulphuric acid. zinc sulphate. hydrogen.
4. Fe + H_2SO_4 = FeSO_4 + H_2
iron. sulphuric acid. ferrous sulphate. hydrogen.
5. Zn + 2HCl = ZnCl_2 + H_2
zinc. hydrochloric acid. zinc chloride. hydrogen.
6. Fe + 2HCl = FeCl_2 + H_2
iron. hydrochloric acid. ferrous chloride. hydrogen.
7. Fe_2 + $4\text{H}_2\text{O}$ = Fe_2O_3 + 4H_2
iron. steam. magnetic oxide of iron. hydrogen.
8. Zn + 2KHO = K_2ZnO_2 + H_2
zinc. caustic potash. potassium zinc oxide. hydrogen.

OXYGEN.

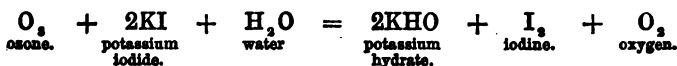
Preparation.

1. $2\text{HgO} = \text{Hg}_2 + \text{O}_2$
mercuric oxide. mercury oxygen.
2. $2\text{KClO}_3 = 2\text{KCl} + 3\text{O}_2$
potassium chlorate. potassium chloride. oxygen.
3. $8\text{MnO}_2 = \text{Mn}_2\text{O}_7 + \text{O}_2$
manganese dioxide. trimanganic tetroxide. oxygen.
4. $\text{MnO}_2 + 2\text{KClO}_3 = 2\text{KCl} + \text{MnO}_2 + 3\text{O}_2$
manganese dioxide. potassic chlorate. potassic chloride. manganese dioxide. oxygen.
5. $2\text{H}_2\text{O} + 2\text{Cl}_2 = 4\text{HCl} + \text{O}_2$
steam. chlorine. hydrochloric acid. oxygen.
6. $2\text{MnO}_2 + 2\text{H}_2\text{SO}_4 = 2\text{MnSO}_4 + 2\text{H}_2\text{O} + \text{O}_2$
manganese dioxide. sulphuric acid. manganese sulphate. water. oxygen.

OZONE.

 O_3 .*Preparation.*

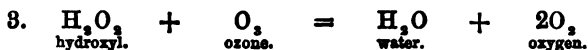
1. By passing a series of electric sparks through air or oxygen.
2. By the action of phosphorus on moist air.
3. By passing an electric current through dilute sulphuric acid.

Test.

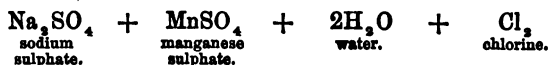
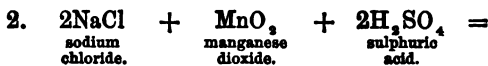
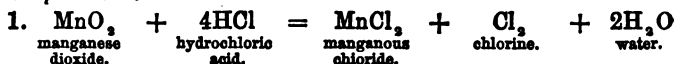
HYDROGEN DIOXIDE.

 H_2O_2 .*Preparation.*

1. $\text{BaO}_2 + \text{H}_2\text{O} + \text{CO}_2 = \text{BaCO}_3 + \text{H}_2\text{O}_2$
barium peroxide. water. carbon dioxide. barium carbonate. hydrogen dioxide.
2. $\text{BaO}_2 + 2\text{HCl} = \text{BaCl}_2 + \text{H}_2\text{O}_2$
barium peroxide. hydrochloric acid. barium chloride. hydrogen dioxide.

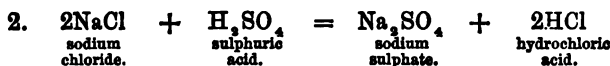
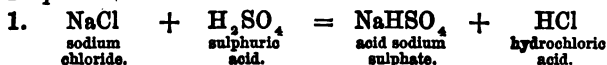
$$1. \quad \begin{array}{ccccccc} \text{PbS} & + & 4\text{H}_2\text{O}_2 & = & \text{PbSO}_4 & + & 4\text{H}_2\text{O} \\ \text{lead} & & \text{hydroxyl} & & \text{lead} & & \text{water.} \\ \text{sulphide.} & & & & \text{sulphate.} & & \end{array}$$


Preparation.

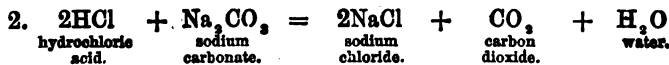

$$1. \text{C}_{10}\text{H}_{16} + 8\text{Cl}_2 = 16\text{HCl} + 10\text{C}$$

turpentine. chlorine hydrochloric acid. carbon

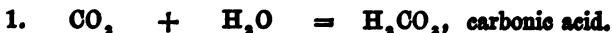
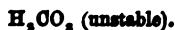
Preparation.


$$1. \quad \text{HCl} + \text{KHO} = \text{KCl} + \text{H}_2\text{O}$$

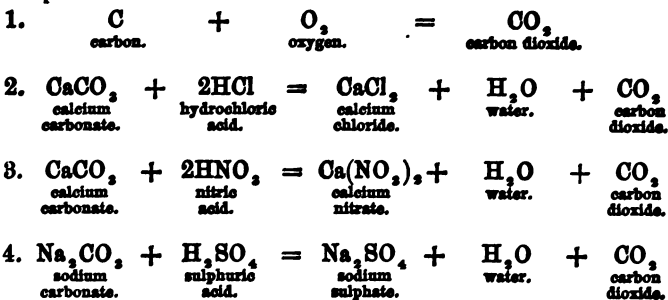
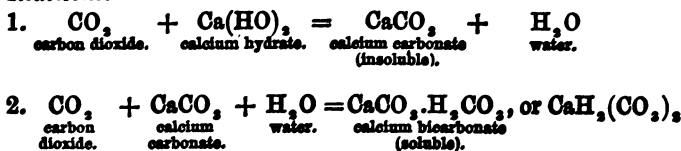
hydrochloric potassium potassium water.
acid. hydrate. chloride.



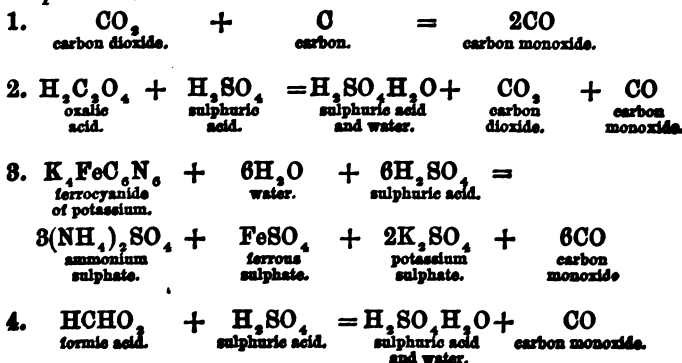
CARBONIC ACID.



CARBON DIOXIDE.

Preparation.*Reactions.*

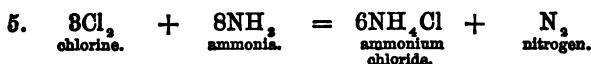
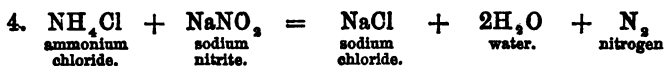
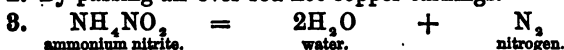
CARBON MONOXIDE.

Preparation.

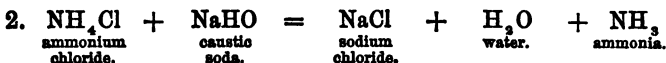
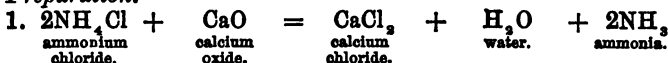
Preparation. NITROGEN.

1. By burning phosphorus in a closed vessel containing air.

2. By passing air over red hot copper turnings.



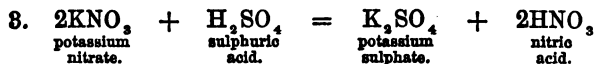
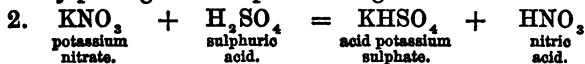
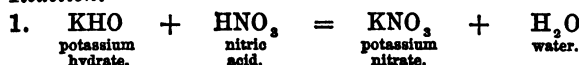
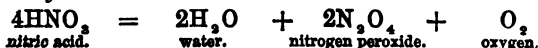
AMMONIA.

Preparation.*Reactions.*

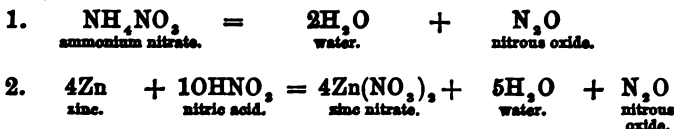
NITRIC ACID.

Preparation.

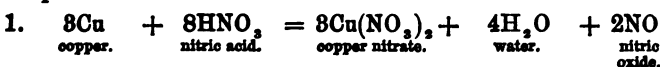
1. By passing electric sparks through moist air.

*Reaction.**Action of heat on nitric acid.*

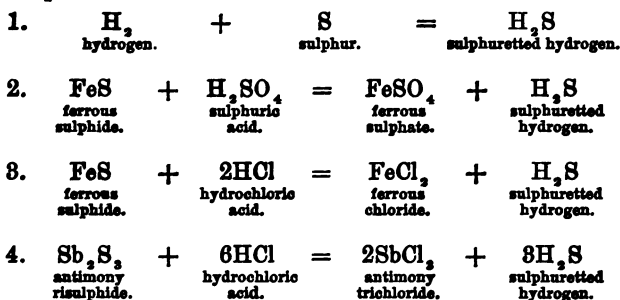
NITROUS OXIDE.

Preparation.

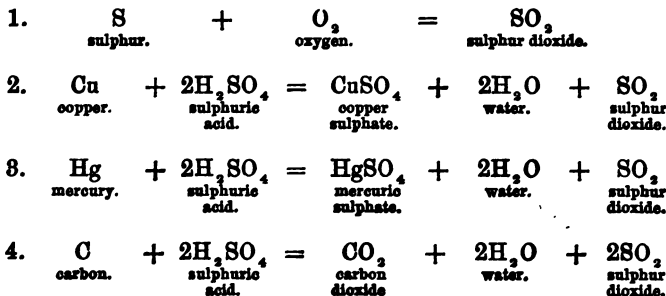
NITRIC OXIDE.

Preparation.

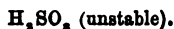
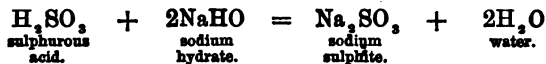
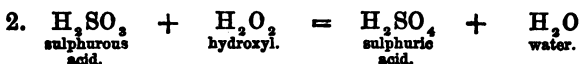
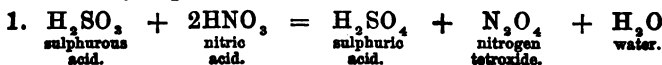
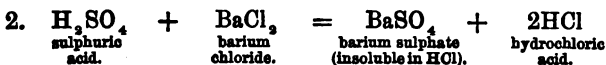
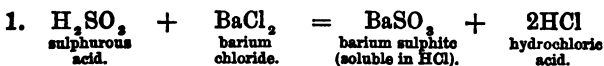
SULPHURETTED HYDROGEN.

Preparation.

SULPHUR DIOXIDE.

Preparation.

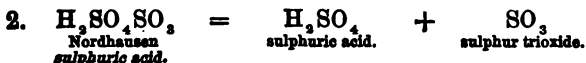
SULPHUROUS ACID.

*Preparation.**Reaction.**Oxidation of sulphurous acid:—**Distinction between sulphurous and sulphuric acids:—*

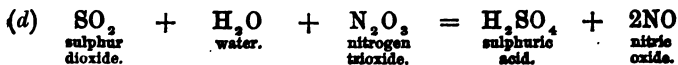
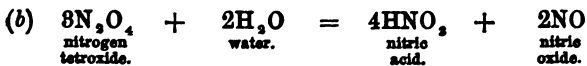
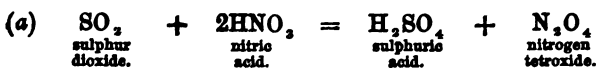
Sulphites are recognised by the characteristic odour of sulphur dioxide which is given off on the addition of a strong acid, thus:—



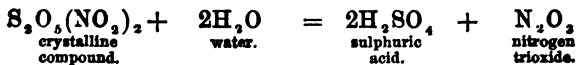
SULPHUR TRIOXIDE.

Preparation.

SULPHURIC ACID.

Preparation.*Manufacture.*

Action of water on the leaden chamber crystals :—



QUESTIONS.

*Questions marked * are taken from the Science and Art
Department Examination Papers.*

*In the examples marked † the gases are supposed to be at the normal
temperature and pressure.*

1. Give the names of the elements represented by the following symbols :—Al, Sb, As, Ba, Bi, B, Br, Cd, Ca, Cr, Co, Cu, F, Au, I, Fe, Pb, Mg, Mn, Hg, Ni, N, P, Pt, H, Se, Ag, Si, Sr, Sn.
2. Give the atomic weights of the following elements :—Ca, C, Cl, Cu, H, Fe, Mn, Hg, N, O, Na, S, Zn.
3. Write the names and symbols of the non-metals.
4. Divide the following elements into metals and non-metals :—Hg, P, Pt, Br, Ba, Ag, Cl, Si, Na.
5. Define the following terms :—Element, compound, mixture. Give three examples of each.
6. The symbol for sulphuric acid is H_2SO_4 . State fully what these symbols indicate.
7. The following formulæ are intended to represent calcium nitrate :— $Ca(NO_3)_2$, $Ca 2NO_3$, $Ca 2(NO)_3$. One of them is wrong ; correct it, and say what difference the mistake would make in the weight.
8. Define atom, molecule, combining weight, and atomic weight.
9. Divide the following list into elements and compounds :—Chalk, charcoal, water, sulphur, lime, sand, iron, lead, diamond.
10. How would you distinguish between a mixture of iron filings and sulphur, and a chemical compound of these two substances ?
11. What is the difference between chemical action and mere physical change ? Give an illustration.
12. A knife blade is dipped into a solution of copper chloride. What is the result and explain it by an equation ?
13. Can the above experiment be repeated any number of times ? Give a reason for your answer.
14. In the experiment (Ques. 12) does the solution gain or lose in weight, and why ?

15. What occurs when mercuric chloride is added to potassium iodide? Give the equation.

16. How would you separate the following mixtures :—

(i.) Sand and salt.

NOTE :—Salt dissolves in water; sand does not.

(ii.) Salt and chalk.

NOTE :—Salt dissolves in water; chalk does not.

(iii.) Sand and chalk.

NOTE :—(a) Chalk is calcium carbonate, and carbonates are decomposed by HCl, or any acid; sand is not acted upon by HCl. Ammonium carbonate added to a solution of calcium chloride gives a precipitate of chalk.

(b) Chalk is soluble in water containing carbon dioxide.

(iv.) Potassium chlorate and potassium chloride.

NOTE :—Both substances dissolve in water, the latter more readily than the former, and on boiling down carefully the chlorate crystallizes out before the chloride.

(v.) Charcoal and sulphur.

NOTE :—Sulphur is dissolved in carbon disulphide; charcoal is unacted upon by it. Carbon disulphide will volatilize.

(vi.) Nitre and sulphur.

NOTE :—Nitre is soluble in water.

(vii.) Sand and ammonium chloride?

NOTE :—Ammonium chloride is volatile, and soluble in water.

17. *A white powder is shaken up with distilled water. How would you ascertain whether any of it dissolves?

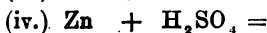
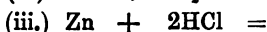
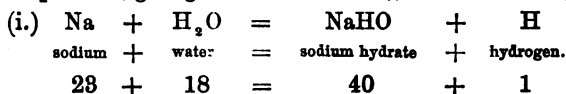
18. *A glass full of dirty water is given to you. Describe, with sketch, the plan you would adopt, first for rendering the water clear; and, secondly, for ascertaining whether the clear water contains any dissolved solid matter.

19. How would you obtain pure water from sea water?

HYDROGEN.

20. Give the atomic weight, molecular weight, specific gravity, and the weight of one litre of hydrogen.

21. Hydrogen may be prepared by the following methods. Complete the equations, giving the names and weights as shown in (i.) :—



22. Sketch and describe the apparatus you would use to obtain hydrogen :—

- (a) From water by the addition of sodium.
- (b) From dilute sulphuric acid by the addition of zinc.
Show carefully the position of the thistle funnel and delivery tubes.
- (c) From water by passing it in the form of steam over red hot iron turnings.

23. Describe fully the experiments you would make to show—

- (a) The extreme lightness of hydrogen.
- (b) That it burns but does not support combustion.
- (c) That when hydrogen burns in air, water is formed.
- (d) That hydrogen diffuses with great readiness.

24. †Give the weights of the following volumes of hydrogen :—
1·5 litre, 10 litres, 5·5 litres.

25. †What volumes will the following weights of hydrogen occupy :—1 gram, ·5 gram, ·25 gram, 2·5 grams, 10 grams?

26. What weight of water can be decomposed by 115 grams of sodium ?

27. †How much sodium will be required with which to obtain 11.16 litres of hydrogen from water ?

28. When potassium was thrown on water 72 grams of water were decomposed. How much caustic potash remained ?

29. How many grams of hydrogen may be obtained by acting on zinc with 196 grams of sulphuric acid ?

30. What weight of hydrogen would be liberated from hydrochloric acid by the addition of 195 grams of zinc ; and how much zinc chloride would remain ?

31. †What measure of hydrogen could be obtained by the addition of 892 grams of sulphuric acid to iron ?

32. †It is required to obtain 2 litres of hydrogen from sulphuric acid. How much zinc must be dissolved in it for this purpose ?

33. †How many litres of hydrogen are produced by passing 216 grams of water, in the form of steam, over red hot iron turnings ?

34. If 216 grams of steam be decomposed by red hot iron turnings, how much magnetic oxide of iron will be formed ?

35. †How much caustic potash must be added to zinc, so that on boiling, 11.16 litres of hydrogen may be liberated ?

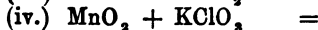
OXYGEN.

36. Give the atomic weight, molecular weight, and the weight of one litre of oxygen.

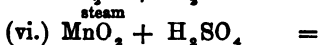
37. What proportion of oxygen, by weight and by volume, is contained in the atmosphere, and in water ?

38. How would you prove by experiment that mercuric oxide is a compound and not an elementary substance ? Sketch and describe the apparatus you would use for this purpose.

39. Oxygen may be prepared by the following methods. Complete the equations, giving the names and weights (compared with H) of each substance :—



(v.) By passing steam and chlorine through a red hot porcelain tube.



40. What are the resulting products when hydrogen, carbon, sulphur, phosphorus, sodium, iron and zinc are burnt in oxygen? Express the reactions by equations.

41. †Give the weights of the following volumes of oxygen:—
1·5 litre, 10 litres, 5·5 litres.

42. †What volumes will the following weights of oxygen occupy:—
1 gram, 16 grams, 24 grams, 100 grams?

NOTE.—Because 1 gram of H occupies 11·16 litres.

∴ 16 grams of O occupy 11·16 litres.

and 1 gram of O occupies $\frac{11·16}{16}$ litres.

= ·6975 litres.

43. How much mercury will remain after obtaining 48 grams of oxygen by heating mercuric oxide?

44. If 783 grams of manganese dioxide be heated, how much oxygen by weight can be obtained?

45. What weight of potassium chlorate is required to produce 144 grams of oxygen?

46. How many grams of oxygen can be obtained from 490 grams of chlorate of potash?

47. What weight of potassium chloride remains after obtaining 8 grams of oxygen from potassium chlorate?

48. †On heating potassium chlorate, exactly 33·48 litres of oxygen were obtained. How much of the salt was used?

49. †If 216 grams of water are decomposed by the electric current, how many litres of oxygen may be collected?

50. †What weight of water was decomposed by electrolysis if the oxygen obtained measured 22·32 litres?

51. Some sulphuric acid was heated with manganese dioxide. It was afterwards found that 1057 grams of manganese sulphate remained. What weight of oxygen was liberated?

52. Steam and chlorine are passed together through a red-hot tube. It is found that 292 grams of hydrochloric acid are produced. What weight of oxygen was liberated?

53. How much oxygen by weight is contained in 100 parts of each of the following substances:—(i.) mercuric oxide, (ii.) water, (iii.) sulphuric acid?

54. After heating some copper in oxygen it was found to have increased in weight 80 grams. What weight of copper oxide was formed?

OZONE.

55. Write the symbol and graphic formula for ozone. Give the molecular weight, also the weight of one litre of ozone.

56. Describe carefully the method you would adopt in order to prepare ozone.

57. Describe the test you would use to detect the presence of ozone.

58. Mention some properties of ozone.

59. Complete the equation :—



60. Ozone is termed *allotropic* oxygen. Explain the meaning of the term *allotropic*.

COMBUSTION.

61. Describe fully the chemistry of the candle flame. How would you show that carbon dioxide and water result from the burning of a candle ?

62. Describe, with sketch, the principle of the Bunsen burner.

63. Matter cannot be destroyed. Describe fully how you would prove this by experiment.

64. Explain the terms "combustion," "combustible gas," "supporter of combustion," and "point of ignition."

65. Hydrogen burns, but does not support combustion ; oxygen does not burn, but supports combustion. When are these statements true, and when not true ?

66. Describe and explain the "Davy lamp."

67. The blowpipe flame consists of two parts, called the *oxidising* flame, and the *de-oxidising* or *reducing* flame. Explain why they are so called.

WATER.

68. Give the formula, molecular weight, density, and weight of one litre of water (steam).

69. You have some spring water and distilled water given to you in separate glasses. How would you decide which was which ?

70. Describe fully an experiment which shows by analysis (splitting up) that water is composed of hydrogen and oxygen in the proportion of two volumes of hydrogen to one volume of oxygen.

71. Describe an experiment which proves by synthesis (putting together) that water is composed of exactly two volumes of hydrogen and one volume of oxygen.

72. It may be shown by experiment that when two volumes of hydrogen unite with one volume of oxygen, *two* volumes of steam at the same temperature and pressure are the result. Describe, with sketch, the apparatus used.

73. Sketch and explain the apparatus used to show the composition of water by weight.

74. †If 100 grams of water are decomposed by an electric current, how many litres of hydrogen and of oxygen will be obtained ?

75. A gas bag contains 144 grams of oxygen. How many grams of hydrogen are required to unite with all the oxygen in order to form water ?

76. †What volume of oxygen is required to unite with 100 grams of hydrogen ?

77. What weight of water must be used to obtain 10 grams of oxygen by electrolysis ?

78. If 20 grams of hydrogen are passed through a heated tube containing an excess of copper oxide, what weight of water would be formed ?

79. Find the quantity of copper remaining when 159 grams of copper oxide have been reduced in a current of hydrogen.

80. If a mixture of 28 c.c. of hydrogen and 14 c.c. of oxygen at 182°C . be exploded, what volume of steam will be formed ?

81. We mix together 30 c.c. of hydrogen, and 18 c.c. of oxygen. If the mixture be exploded at 150°C , what volume of gas will remain, and what will it be composed of ?

82. In the previous question substitute 25 c.c. of hydrogen, and 15 c.c. of oxygen.

83. †Hydrogen burns at a jet in oxygen till 10 litres of the hydrogen are consumed. What volume of oxygen took part in the burning, and what weight of water resulted ?

84. We require sufficient oxygen to unite with all the hydrogen in 892 grams of sulphuric acid. How much mercuric oxide must be decomposed in order to obtain the exact quantity of oxygen ?

PEROXIDES.

85. How may red lead (Pb_3O_4) be obtained from the yellow oxide of lead (PbO) ?

86. Describe by experiment the action of a strong heat on red lead.

87. It has been proposed to use barium oxide as a means of obtaining oxygen on a large scale. Describe experiments explaining this use of barium oxide.

88. How is hydrogen dioxide (hydroxyl) prepared ? Give the equation.

89. Describe a test used to detect the presence of hydrogen dioxide.

90. What are the properties of hydrogen dioxide ?

91. Explain the action of hydrogen dioxide in renovating old oil paintings.

92. Explain the meaning of the terms "oxidising" and "reducing" agents, and give examples.

93. Carbon dioxide is passed into water to which 388 grams of barium dioxide are added, until the whole of the latter is converted into carbonate with liberation of oxygen. What weight of hydrogen peroxide is formed ?

94. What weight of barium dioxide will be required in order to supply sufficient oxygen to produce 102 grams of hydrogen dioxide ?

95. Hydrochloric acid is added to barium dioxide with the formation of 416 grams of barium chloride. What weight of hydrochloric acid was used, and what is the amount of hydrogen dioxide formed ?

96. Find the weight of oxygen which may be obtained by heating 306 grams of hydrogen dioxide.

97. What percentage of oxygen by weight is contained in hydrogen dioxide ?

CHLORINE.

98. What is the molecular weight, density, specific gravity, and the weight of one litre of chlorine ?

99. Mention two ready methods of obtaining chlorine (i.) from hydrogen chloride, (ii.) from sodium chloride, and give the equations. Sketch and describe the apparatus used.

100. Chlorine may be collected by downward displacement, but not over water or mercury. Explain the reason in each case.

101. The following substances are introduced into jars of chlorine gas. State the results and give equations :—

- (i.) A lighted taper.
- (ii.) Finely powdered antimony.
- (iii.) A piece of phosphorus.
- (iv.) A piece of sodium.
- (v.) A solution of litmus.
- (vi.) A burning jet of hydrogen.
- (vii.) A piece of filter paper, or tow, soaked in turpentine ($C_{10}H_{16}$).

102. Describe the apparatus you would use for decomposing steam by chlorine.

103. Explain the bleaching action of chlorine.

104. Chlorine water is decomposed :—

- (i.) When exposed to sunlight.
- (ii.) On addition of sulphuretted hydrogen water.

Explain this.

105. †Calculate the weights of the following volumes of chlorine:—1·5 litre; 2·5 litres; 10 litres.

106. †What volumes will the following weights of chlorine occupy :—35·5 grams, 71 grams, 284 grams?

107. It is required to obtain 35·5 grams of chlorine. How much manganese dioxide must be heated with hydrochloric acid in order to produce that quantity?

108. Find the quantity of hydrochloric used, and of manganous chloride left, in the previous question.

109. †What weight of manganese dioxide must be heated with hydrochloric acid in order to obtain 88·48 litres of chlorine?

110. †What weight and what volume of chlorine can be obtained from 58·5 grams of common salt, by the addition of sulphuric acid and manganese dioxide?

111. †What weight of sodium chloride must be used in order to obtain from it exactly 1 litre of chlorine?

HYDROCHLORIC ACID.

112. †What is the molecular weight, density, specific gravity, and the weight of one litre of hydrochloric acid gas ?

118. How may hydrochloric acid be prepared ? Describe the apparatus used and give equations.

114. You are supplied with the following apparatus :—three jars of hydrochloric acid gas, a taper, a basin of water, a solution of litmus, and a solution of sodium hydrate. What properties of hydrochloric acid gas would you illustrate with these ?

115. *Hydrochloric acid is stated to be composed of equal volumes of chlorine and hydrogen united without condensation. How would you prove experimentally that this is the case (a) by analysis ; (b) by synthesis ?

116. What happens when gaseous hydrochloric acid is passed over manganese dioxide ?

117. You have given to you some sulphuric acid, manganese dioxide, and silver nitrate. How would you proceed with these to test for the presence of chlorine ?

118. †Give the weights of the following volumes of hydrochloric gas :—2 litres, 3·5 litres, of 5·5 litres.

119. †What volumes will the following weights of hydrochloric acid gas occupy :—18·25 grams, 2 grams, 5 grams ?

120. It is required to obtain 78 grams of hydrochloric acid from sodium chloride by the addition of sulphuric acid. What weight of sodium chloride must be used for this purpose ?

121. †How much sulphuric acid by weight must be added to sodium chloride in order to produce at a gentle heat 11·16 litres of hydrochloric acid gas ? Also, what weight of acid sodium sulphate remains ?

122. †A certain quantity of hydrogen and chlorine were mixed and exploded, forming 78 grams of hydrochloric acid. What were the quantities, by weight and by volume, of both hydrogen and chlorine ?

123. What weight of sodium chloride would be formed after passing 146 grams of hydrochloric acid gas into a solution of sodium hydrate ?

CARBON.

124. *Describe the allotropic forms of carbon. How would you prove that these different substances consist of the same element?

125. In what way are putrid gases rendered harmless by charcoal?

126. A piece of pure carbon was burnt completely in a current of air. The resulting carbon dioxide weighed 11 grams. What was the weight of the carbon?

CARBON DIOXIDE.

127. You are required to prove that carbon dioxide (carbonic acid gas) is formed when wood is burnt. Describe and sketch the apparatus you would use.

128. What is the action of hydrochloric acid on marble? Sketch and describe the apparatus you would use to collect the gaseous product; also give the equation. Would you prefer to use sulphuric acid in this experiment, and why?

129. You are required to explain the following results, and in each case to give an equation:—

(i.) Carbon dioxide is passed into lime water, and the lime water is turned milky.

(ii.) The carbon dioxide is passed for some time, when the milky solution at length becomes clear.

(iii.) The clear solution in (ii.) is divided into two parts (a) and (b).

(a) is boiled, when it again becomes milky.

(b) lime water is added, and it again becomes milky.

(iv.) To one of the milky solutions hydrochloric acid is added; the solution effervesces and at the same time becomes clear.

130. Describe as many experiments as you can to show that carbon dioxide is heavier than air.

131. What is *hard* water? How may it be softened? Also explain the formation of a crust on the inside of a tea kettle.

132. The following are introduced into jars of CO_2 :—

(i.) A lighted taper.

(ii.) A piece of brilliantly burning magnesium ribbon.

(iii.) A solution of blue litmus, which is afterwards boiled.

What is the result in each case?

183. What occurs when carbon dioxide is passed into the following :—

- (i.) Baryta water.
- (ii.) Solution of caustic soda.
- (iii.) Solution of caustic potash ?

Give equations.

184. How would you prove the presence of carbon dioxide, and moisture in the breath ?

185. †Calculate the weights in grams of the following volumes of carbon dioxide :—2·5 litres, 5·5 litres, 10 litres.

186. †What volumes will the following weights of carbon dioxide occupy :—44 grams, 1 gram, 5·5 grams ?

187. †What quantity of oxygen by weight and by volume is required to unite with pure carbon to form 66 grams of carbon dioxide ?

188. †After burning charcoal in air it was found that 42 litres of carbon dioxide had been formed. What volume of air was used ?

189. *If 12 grams of pure carbon be completely burnt in the oxygen which is obtained by decomposing 122·5 grams of potassic chlorate, what is the weight of the product of the combustion formed, and what is the weight, if any, of the oxygen remaining ?

140. *How many litres of carbon dioxide (carbonic anhydride) will be formed by the complete combustion of 12 grams of carbon, and how many litres of oxygen will be required ?

141. †You burn 6 grams of carbon in 12 litres of oxygen. What volume of carbon dioxide is formed, and what volume of oxygen, if any, remains ?

142. How much lime (CaO) is contained in 100 grams of calcium carbonate ?

143. Some hydrochloric acid was added to some marble, when 182 grams of carbon dioxide were liberated. What weights of both hydrochloric acid and marble took part in the reaction ?

144. †What weight of marble when acted upon by an acid would supply us with exactly one litre of carbon dioxide ?

145. What weight of water must be decomposed in order to supply sufficient oxygen to form, with pure carbon, 44 grams of carbon dioxide ?

146. An excess of hydrochloric acid was added (a) to 100 grams of calcium carbonate, (b) to 100 grams of sodium carbonate. Which substance supplied the larger quantity of carbon dioxide? and give the difference in grams.

147. Which liberates the larger quantity of carbon dioxide from an excess of carbonate—100 grams of sulphuric acid, or 100 grams of hydrochloric acid? also give the difference in grams.

148. Carbon dioxide is passed into a solution of calcium hydrate (lime water) and the precipitate formed is collected, dried, and weighed. It weighs 5 grams. How much carbon dioxide does it contain?

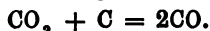
149. †What weight of sodium carbonate is formed after passing 22.82 litres of carbon dioxide into an excess of sodium hydrate solution?

CARBON MONOXIDE.

150. †What is the molecular weight, density, specific gravity, and the weight of one litre of carbon monoxide?

151. Explain the appearance of blue flame on the top of an ordinary coal fire.

152. Explain the following equation, and sketch and describe the apparatus you would use to bring about the result:—



153. *Oxalic acid is heated with sulphuric acid. How would you experimentally prove that the gases evolved are carbonic anhydride (CO_2) and carbonic oxide (CO)?

154. How would you obtain carbon monoxide from potassium ferrocyanide? Sketch the apparatus and give the equation.

155. *Lime water is shaken up in a jar filled with carbonic oxide; what occurs? The carbonic oxide is then inflamed, and the lime water once more shaken up with the contents of the jar. What now takes place?

156. What are the principal properties of carbon monoxide?

157. †Calculate the weights in grams of the following volumes of carbon monoxide (carbonic oxide):—2.5 litres, 5.5 litres, 7.5 litres.

158. †Give the volumes, in litres, which the following weights of carbon monoxide occupy:—28 grams, 7 grams, 50 grams.

159. †A certain volume of carbon dioxide weighs 396 grams, what is the weight of the same volume of carbon monoxide, and of the *same volume* of hydrogen?

160. What percentage of carbon by weight is contained in carbon monoxide, and in carbon dioxide?

161. †If 5·58 litres of carbon dioxide be passed through a tube containing red hot charcoal, what volume and weight of carbon monoxide will be produced?

162. †A litre of carbon monoxide is burnt in air. What volume and what weight of carbon dioxide is formed?

163. *How many volumes of carbon dioxide (carbonic acid gas) will be formed when a mixture of four volumes of carbon monoxide (carbonic oxide), and four volumes of oxygen is burnt? State what volume of oxygen, if any, remains uncombined.

164. If 20 grams of carbon are heated in the presence of 44 grams of carbon dioxide, what weight of carbon monoxide is formed, and what weight, if any, of carbon remains?

165. †What weight of oxalic acid must be used, so that on heating with sulphuric acid, 11·16 litres of carbon monoxide may be collected?

166. †What volume and what weight of carbon dioxide, and carbon monoxide, can be obtained from 180 grams of oxalic acid?

167. †What weight and what volume of carbon monoxide can be obtained by heating 92 grams of potassium ferrocyanide with sulphuric acid?

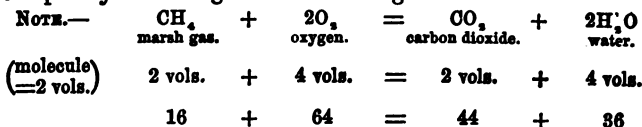
168. How much yellow prussiate of potash must be used from which to obtain 1 gram of carbon monoxide?

MARSH GAS.

169. †Calculate the weights of the following volumes of marsh gas (CH_4):—1 litre, 5·5 litres, 8·5 litres.

170. †Give the volumes which the following weights of marsh gas occupy:—8 grams, 1 gram, 5·5 grams.

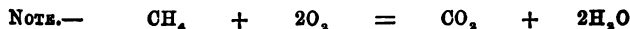
171. †What weight and what volume of oxygen will be required to completely burn 82 grams of marsh gas?



Hence 64 grams of oxygen are required to completely burn 16 grams of marsh gas. Remember also, 16 grams of oxygen occupy 11·16 litres.

172. †What weight of carbon dioxide and water are formed by the combustion of 12 grams of marsh gas in oxygen or air?

173. †What volume of air is required to supply sufficient oxygen to burn 10·5 litres of marsh gas?



$$2 \text{ vols.} + 4 \text{ vols.} = 2 \text{ vols.} + 4 \text{ vols.}$$

$$\text{or } 1 \text{ vol.} + 2 \text{ vols.} = 1 \text{ vol.} + 2 \text{ vols.}$$

$$\left(\begin{array}{l} \text{multiplying} \\ \text{throughout} \\ \text{by } 10\cdot5 \end{array} \right) 10\cdot5 \text{ vols.} + 21 \text{ vols.} = 10\cdot5 \text{ vols.} + 21 \text{ vols.}$$

From the above we see the volume of oxygen required to burn 10·5 litres of marsh gas, viz.:—21 litres. The question now is—How many litres of air supply 21 litres of oxygen?

174. †We burn 42 litres of marsh gas (*a*) in oxygen, (*b*) in air. What volume of each was used?

NITROGEN.

175. How would you proceed to obtain nitrogen—

- (i.) From air by the aid of phosphorus?
- (ii.) From air by the aid of copper?
- (iii.) From ammonia? What precaution must be taken?
- (iv.) From ammonium nitrite? Give equation.

176. How is the ammonium nitrite prepared?

177. What are the properties of nitrogen?

178. What is the result of passing electric sparks through moist air in a closed vessel?

179. †Calculate the weights of the following volumes of nitrogen:—2·5 litres, 5·5 litres, 8·5 litres.

180. †Give the volumes which the following weights of nitrogen will occupy:—1 gram, 14 grams, 21 grams.

181. †What weight of ammonium nitrite must be heated so as to obtain exactly 1 litre of nitrogen?

182. †What volume of nitrogen can be obtained by heating 128 grams of ammonium nitrite?

THE ATMOSPHERE.

183. What is the composition of the atmosphere? Also give, as nearly as you can, the proportion by weight and by volume of its two most important constituents.

184. How may the composition of the atmosphere be determined (i.) by weight, and (ii.) by volume?

185. Is the atmosphere a mixture or a compound? Give as many reasons as you can for your answer.

186. What effect is produced on the atmosphere—

(i.) By the burning of coal fires?

(ii.) By the breathing of animals?

(iii.) By the action of plants?

Describe the experiments you would make in proof of your answers.

187. *Describe three experiments, the best you can think of, one is to prove that the air contains oxygen, another that it contains carbon dioxide, and the third, that it contains aqueous vapour.

188. †What are the weights of the following volumes of air :—
5 litres, 8·5 litres, 10 litres?

NOTE.—Air is 14·42 times as heavy as hydrogen.

1 litre of H = ·0896 grams

∴ 1 litre of air = (·0896 × 14·42) grams

= 1·292 grams

and 5 litres „ = (1·292 × 5) grams

= 6·46 grams

189. †What volumes will the following weights of air occupy :—
28·84 grams, 7·21 grams, 100 grams?

NOTE.—1 gram of H occupies 11·16 litres.

∴ 14·42 grams of air occupy 11·16 litres.

and 1 gram of air occupies $\frac{11·16}{14·42}$ litres.

∴ 28·84 grams.....occupy $\frac{11·16 \times 28·84}{14·42}$ litres
= 22·32 litres

190. †What volume of nitrogen should be added to 63 volumes of oxygen to form a mixture similar to air? (Air contains 21 per cent of oxygen and 79 per cent of nitrogen by volume.)

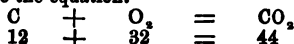
191. A bottle contains 16 grams of a gas which is proved to be oxygen. What weight of air at the same temperature and pressure would three such bottles contain?

192. †We have two litre flasks ; one contains equal volumes of nitrogen and oxygen, the other is filled with air. Which contains the heavier mixture, and give the difference of weight in milligrams ?

$$\begin{array}{rcl} \text{NOTE.—1 litre of H} & = & \cdot 0896 \text{ grams} \\ 1 \text{ ,, ,, N} & = & (\cdot 0896 \times 14) \text{ grams} \\ \frac{1}{2} \text{ ,, ,, N} & = & \frac{\cdot 0896 \times 14}{2} \text{ grams} \\ \frac{1}{2} \text{ ,, ,, O} & = & \frac{\cdot 0896 \times 16}{2} \text{ grams} \\ 1 \text{ ,, ,, air} & = & (14 \cdot 42 \times \cdot 0896) \text{ grams} \end{array}$$

193. Air contains 23 per cent of its weight of oxygen. How many grams of carbon are required to unite with all the oxygen in 100 grams of air, with the formation of carbon dioxide ?

NOTE.—First write the equation.



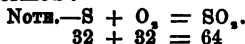
By the question every 100 grams of air contain 23 grams of oxygen ; we therefore require to know how many grams of carbon unite with the 23 grams of oxygen to form carbon dioxide.

By the equation, we see that 12 grams of carbon unite with 32 grams of oxygen, hence

$$\begin{array}{ccccc} \text{oxygen} & & \text{oxygen} & & \text{carbon} \\ 32 & : & 23 & :: & 12 : x \end{array}$$

194. How many grams of phosphorus are required to unite with all the oxygen in 100 grams of air, with the formation of phosphorus pentoxide ?

195. Some sulphur is burnt in 100 grams of air. What weight of sulphur dioxide is formed ?



By the equation, 64 grams of sulphur dioxide are formed from 32 grams of oxygen.

196. What weight of hydrogen is required to unite with all the oxygen in 100 grams of air, to form water ?

197. †What volume of hydrogen is required to unite with all the oxygen in 100 litres of air, so as to form water ?

198. †After passing air in excess over red-hot charcoal it was found that 21 litres of CO_2 were formed. What volume of air has been deprived of its oxygen ?

199. If 100 grams of air are passed slowly over red-hot copper, what weight of copper oxide will be formed ? and what weight of nitrogen may be collected ?

200. How many kilograms of air are required to furnish us with 100 kilograms of oxygen ?

AMMONIA.

201. What are the molecular and graphic formulæ for ammonia? Also give its molecular weight, density, specific gravity, and the weight of one litre of ammonia.

202. How would you prepare and collect ammonia? Sketch the apparatus you would use, and give the equation.

203. Describe and explain the results in the following experiments. You have given to you four jars of ammonia; the first is placed with its mouth under water coloured with red litmus solution; the second is placed over mercury, and a piece of charcoal inserted; in the third a lighted taper is immersed; the fourth is placed mouth to mouth with a jar of hydrochloric acid.

204. Sketch the apparatus used, and explain the result in the following cases:—

- (i.) Ammonia gas is passed through a red hot iron tube, and a light is applied at the end of the tube.
- (ii.) Ammonia and oxygen are passed together through a glass tube, and a light is applied at the end of the tube.
- (iii.) Well dried ammonia gas is passed over heated potassium.

205. Explain the difference between ammonia and ammonium.

206. You have two beakers, each containing a solution of ammonia. Into one carbon dioxide is passed for some time; into the other sulphuretted hydrogen is passed. What is the result in each case, and give equations?

207. Ammonia forms salts with acids. Give equations showing how the chloride, nitrate, and sulphate of ammonium are formed.

208. What is the action of heat on the chloride, nitrate, and phosphate of ammonia?

209. †Calculate the weights of the following volumes of ammonia:—1.5 litre, 5.5 litres, 10 litres.

210. †Give the volumes which the following weights of ammonia will occupy:—4.25 grams, 17 grams, 25 grams.

211. †A certain volume of air weighs 28.84 grams, what is the weight of the same volume of ammonia?

212. What is the percentage composition of ammonia, by weight?

213. †What volume of nitrogen and hydrogen can be obtained by passing electric sparks through 1 litre of dry ammonia gas? and what volume of oxygen will be required to unite with all the hydrogen thus obtained, to form water?

214. †Twenty-eight c.c. of ammonia gas are decomposed by electric sparks, and one hundred c.c. of air are then added to the mixture. Electric sparks are again passed. What is the residue composed of, and how many c.c. of the gas, or gases, remain?

215. †What volume of ammonia may be obtained by heating 214 grams of ammonium chloride with lime?

216. †What weight of ammonium chloride shall we require from which to obtain 1 litre of ammonia?

217. We pass 68 grams of ammonia through a red hot tube, and burn the gas at the end. Name and give the weights of the products.

218. What weight of ammonia will neutralize 146 grams of hydrochloric acid?

219. †What volume of ammonia neutralizes 98 grams of sulphuric acid; and what is the weight of the product formed?

220. Which requires the larger amount of ammonia in order to neutralize it, 100 grams of hydrochloric acid, or 100 grams of nitric acid? and give the difference in grams.

NITRIC ACID.

221. What is the molecular weight, specific gravity, and the weight of one litre of nitric acid vapour?

222. Describe in detail the preparation of nitric acid. Sketch the apparatus, and give the equation.

223. What are the properties of nitric acid?

224. What is the result of adding nitric acid—

(i.) To a solution of caustic soda?

(ii.) To a solution of sodium carbonate?

Give equations.

225. What is "aqua regia"? Describe fully its action on gold.

226. Nitric acid and nitrates are powerful oxidising agents. Describe any experiments you know in proof of this.

227. Describe carefully how you would test for the presence of nitric acid, or nitrates.

228. †Calculate the weights of the following volumes of nitric acid vapour:—2.5 litres, 5.5 litres, 10 litres.

229. †Give the volumes which the following weights of nitric acid vapour occupy:—81·5 grams, 68 grams, 1 gram.

230. What weight of nitric acid can be obtained by heating 25½ pounds of saltpetre with sulphuric acid?

231. What weight of sulphuric acid must be added to nitre, so that on heating in a glass vessel, 15·75 grams of nitric acid may be obtained?

232. What weight of potassium is contained in 100 grams of acid potassium sulphate?

233. Thirty-one and a half grams of nitric acid are neutralised with caustic soda. How much caustic soda was required? Give the name and weight of the product.

NITROGEN MONOXIDE OR NITROUS OXIDE.

234. †Give the molecular weight, specific gravity, and the weight of one litre of nitrogen monoxide.

235. What changes take place when zinc dissolves in dilute nitric acid? Draw the apparatus you would use for this experiment, and give the equation.

236. Carbonate of ammonia, and nitric acid are given to you. How would you prepare nitrous oxide from them? Sketch the apparatus you would use, and give equations.

237. Write two equations comparing the results of heating ammonium nitrite and ammonium nitrate.

238. Describe and explain the results which take place on introducing the following substances into jars of nitrous oxide gas:—

- (i.) A glowing chip.
- (ii.) A piece of sulphur burning feebly.
- (iii.) A piece of sulphur heated strongly.
- (iv.) A piece of ignited phosphorus.

239. Why is nitrous oxide called “laughing gas”?

240. †Find the weights of the following volumes of nitrous oxide:—2·5 litres, 7·5 litres, 10 litres.

241. †Give the volumes which the following weights of nitrous oxide occupy:—22 grams, 68 grams, 1 gram.

242. A certain volume of air weighs 1 gram. What is the weight of the same volume of nitrous oxide at the same temperature and pressure?

243. †How many litres of nitrogen are contained in 7 litres of nitrous oxide ?

244. †If 10 litres of nitrous oxide gas be decomposed into its constituent elements, what alteration of volume will take place ?

245. †What weight of oxygen is contained in 5 litres of nitrogen monoxide ?

246. If three grams of carbon are burnt in an excess of nitrous oxide, what weight of carbon dioxide is formed, and what weight of nitrous oxide is decomposed ?

247. †If 4 grams of sulphur are burnt in an excess of nitrous oxide, what weight of sulphur dioxide is formed, and what volume of free nitrogen remains ?

248. †What weight of carbon is required to unite with all the oxygen in 11.16 litres of nitrous oxide ?

249. What weight of nitrous oxide is required to supply 62 grams of phosphorus with sufficient oxygen to form phosphorus pentoxide ?

250. †If 11.16 litres of nitrous oxide be mixed with excess of hydrogen, and an electric spark be passed through, what volume of hydrogen takes part in the explosion, and what weight of water is formed ?

251. †What weight of water will supply sufficient hydrogen to unite with all the oxygen in 5.58 litres of nitrous oxide ?

252. †What weight and what volume of nitrogen remains after exploding 11.16 litres of nitrous oxide with excess of hydrogen ?

253. What weight of nitrous oxide may be obtained by heating 10 grams of ammonium nitrate ?

254. †What weight of ammonium nitrate must be decomposed by heat, so as to produce exactly 1 litre of nitrous oxide ?

255. †If 65 grams of zinc are heated with an excess of dilute nitric acid, what weight and what volume of nitrous oxide should be produced ?

256. What weight of zinc is required to act on dilute nitric acid, so as to produce 33.48 litres of nitrous oxide ?

257. How many grams of nitric acid are acted upon by zinc when 2 grams of nitrous oxide gas are liberated ?

NITROGEN DIOXIDE OR NITRIC OXIDE.

258. †What is the molecular weight, density, specific gravity, and the weight of one litre of nitric oxide?

259. How may nitric oxide be prepared? Sketch and describe the apparatus and give the equation.

260. Nitric oxide is colourless; but in the generating flask the gas which comes off first is of a dense red colour. Explain the cause of this colour.

261. You have four jars of nitric oxide, into which are introduced the following:—

- (i.) A lighted taper.
- (ii.) A piece of brightly burning sulphur.
- (iii.) A piece of feebly burning phosphorus.
- (iv.) A piece of phosphorus more strongly heated.

Compare the results of these experiments with a similar treatment of nitrous oxide.

262. A jar of nitric oxide is inverted in a small pneumatic trough containing water coloured with blue litmus. A few bubbles of oxygen, or air, are now introduced into the jar. Describe fully the result.

263. The above experiment (Ques. 262) suggests a method of distinguishing between oxygen and nitrogen monoxide. Describe how you would apply this test.

264. †Calculate the weights of the following volumes of nitric oxide:—11·16 litres, 2·79 litres, 1 litre.

265. †Give the volumes of the following weights of nitric oxide:—80 grams, 7·5 grams, 100 grams.

266. †A certain volume of air weighs 28·84 grams. What is the weight of the same volume of nitric oxide?

267. What percentage of oxygen by weight is contained in nitric oxide?

268. What weight of phosphorus, when strongly ignited, will unite with all the oxygen in 150 grams of nitric oxide to form phosphorus pentoxide?

269. †After burning phosphorus in nitric oxide, 14·2 grams of phosphorus pentoxide are formed. What weight of nitric oxide was decomposed, and what volume of nitrogen remained?

270. †What volume of nitric oxide is required in order to supply sufficient oxygen to completely burn 62 grams of phosphorus ?

271. †What weight of nitric acid should be added to copper in order to obtain 11.16 litres of nitric oxide ?

272. †If 22.32 litres of nitric oxide be obtained by addition of nitric acid to copper, what was the weight of the copper used, and of the copper nitrate formed ?

273. What weight of nitric oxide may be obtained by adding an excess of nitric acid to one gram of copper ?

SULPHUR.

274. Give the molecular weight of sulphur, also the weight of one litre of sulphur vapour, at 1000°C. , compared with hydrogen at the same temperature.

275. Name the allotropic forms of sulphur.

276. Describe fully how you would prepare the different varieties of sulphur.

277. *Describe the changes produced in brimstone by the continued application of heat.

278. Sulphur is *dimorphous*. We have also *amorphous* sulphur. Explain the meaning of these terms as applied to sulphur.

279. Draw up a table showing the analogy, or likeness, between oxygen compounds and sulphur compounds.

280. *If I burn a piece of sulphur in a bottle filled with air, and another piece in a bottle filled with oxygen, what shall I find in each of the bottles after the combustion is finished ? Give equations.

SULPHURETTED HYDROGEN.

281. †What is the molecular weight, density, and the weight of one litre of sulphuretted hydrogen ?

282. How is sulphuretted hydrogen commonly prepared ? Sketch and describe the apparatus, and give an equation.

283. How may sulphuretted hydrogen be obtained from antimony trisulphide ? Give the equation.

284. Hydrogen sulphide burns at a jet. What are the products of combustion, and how would you prove your answer ?

285. The following experiments are performed with three jars of hydrogen sulphide :—A lighted taper is placed in the first ; a solution of litmus in the second ; the third is placed mouth to mouth with a jar of chlorine. Describe fully the result in each case.

286. *Describe what you consider the best experimental proof that sulphur is a constituent of sulphuretted hydrogen.

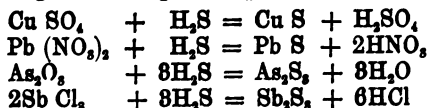
287. You have a solution of lead acetate and a piece of filter paper ; how would you proceed with these to discover the presence of hydrogen sulphide ?

288. A solution of H_2S is decomposed on standing for some time. Explain the reason and give an equation.

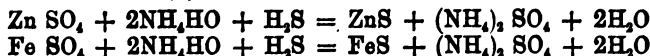
289. You have given to you separate solutions of the following, each of which, except the lead, has been made acid with HCl ; copper sulphate, lead nitrate, arsenious acid, antimony chloride, zinc sulphate, ferrous sulphate, and magnesium sulphate.

(i.) Say what will be the result [of passing H_2S through these solutions, and give equations showing how the precipitated sulphides are obtained.

The following are the required equations :—



(ii.) How shall we proceed to obtain a further precipitation of sulphides in those solutions which remain clear in (i.) ? Give equations.



NOTE.—The NH_4HO neutralises the acid present in the solutions ; also that formed by the action of H_2S on $FeSO_4$. Acids decompose FeS , and so prevent its precipitation.

(iii.) Show how, by the aid of H_2S and an acid and alkali, we may separate solutions of the metals—copper, lead, arsenic, antimony, zinc, iron, and magnesium into three groups.

290. †Calculate the weights of the following volumes of sulphuretted hydrogen :—2·5 litres, 5·5 litres, 10 litres.

291. †Give the volumes which the following weights of sulphuretted hydrogen occupy :—34 grams, 8·5 grams, 1 gram.

292. †A certain volume of sulphuretted hydrogen weighs 51 grams. What is the weight of the same volume of air at the same temperature and pressure ?

293. †What weight and what volume of hydrogen are contained in 102 grams of sulphuretted hydrogen ?

294. What is the percentage composition, by weight, of ferrous sulphide ?

295. †What weight and what volume of sulphuretted hydrogen are obtained by the addition of sulphuric acid to 22 grams of ferrous sulphide ?

296. †What weight of ferrous sulphide, on addition of an excess of sulphuric acid, will liberate 2·79 litres of sulphuretted hydrogen ?

297. †If 66·96 litres of sulphuretted hydrogen were obtained from ferrous sulphide and sulphuric acid, what weight of sulphuric acid was used ?

298. †What weight of hydrochloric acid should be added to ferrous sulphide in order to liberate 5·58 litres of sulphuretted hydrogen ?

299. †What volume of oxygen is required for the complete combustion of 11·16 litres of sulphuretted hydrogen ?

300. On burning sulphuretted hydrogen in air 2·25 grams of water were formed. What weight of H_2S was burnt ?

301. †What weight of sulphur dioxide and water are formed, after the complete combustion of 5·58 litres of sulphuretted hydrogen in oxygen ?

302. Sulphuretted hydrogen gas is passed into a solution of copper sulphate till all the copper is precipitated as sulphide. The weight of copper sulphate present was 6·88 grams. What is the weight of the precipitated sulphide ?

303. What weight of zinc chloride should be dissolved in water, so that on passing an excess of sulphuretted hydrogen through the solution 19·4 grams of zinc sulphide may be obtained ?

304. All the lead in 65 grams of lead acetate was precipitated as sulphide. What weight of sulphuretted hydrogen took part in the reaction ?

305. What weight of chlorine is required to decompose 8·6 grams of sulphuretted hydrogen ? Also give the names and weights of the products.

SULPHUR DIOXIDE AND SULPHUROUS ACID.

806. What is the molecular weight, density, specific gravity, and the weight of one litre of sulphur dioxide ?

807. How may sulphur dioxide be obtained from sulphuric acid ? Give the equation, and sketch the apparatus used.

808. What is the action of sulphuric acid on charcoal ? Give the equation, and say how the gaseous products may be separated.

809. How may sulphur dioxide be obtained in the liquid form ? Sketch the apparatus you would use.

810. What properties of sulphur dioxide could you show with the following apparatus:—Three jars of SO_2 ; a taper ; a trough of water ; a solution of litmus ; a jar of H_2S ?

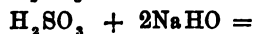
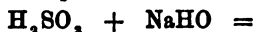
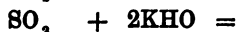
811. Sulphur dioxide is passed into distilled water. Is the result a mixture, or a compound ; also, what is the action on the solution (i.) of heat, (ii.) of long exposure to the air ?

812. Write as many equations as you can illustrating the reducing action of sulphur dioxide. ✓

813. Explain the bleaching action of sulphur dioxide, and show in what way its action differs from that of chlorine.

814. Sulphur dioxide is employed in the arts as an “ antichlor.” Explain this use of the gas, and give an equation.

815. Complete the following equations, giving the name to each formula :—



816. †Calculate the weights of the following volumes of sulphur dioxide :—11·16 litres, 1 litre, 5·58 litres.

817. †Give the volumes occupied by the following weights of sulphur dioxide :—1·5 gram, 8 grams, 10·5 grams.

818. †A certain volume of sulphur dioxide gas weighs 16 grams. What is the weight of the same volume of air, and of carbon dioxide, at the same temperature and pressure ?

819. †One gram of sulphur is burnt in an excess of oxygen. What weight of oxygen unites with the sulphur, and what volume of sulphur dioxide is formed ?

820. †What weight of copper turnings must be heated with strong sulphuric acid, so as to obtain 5.58 litres of sulphur dioxide?

821. How much sulphuric acid must be heated with copper so as to produce 8 grams of sulphur dioxide?

822. What weight of sulphur dioxide can be obtained by heating an excess of strong sulphuric acid with 2.54 grams of copper?

823. If 20 grams of mercury be heated with an excess of strong sulphuric acid, what weight of sulphur dioxide and of mercuric sulphate will be produced?

824. †Six grams of charcoal are heated with an excess of sulphuric acid. What are the weights and volumes of the gaseous products?

825. †What weight of water is required to unite with 1.116 litres of sulphur dioxide to form sulphurous acid?

826. †If 88.48 litres of sulphur dioxide be absorbed by distilled water, what weight of sulphurous acid is formed?

827. †Forty-one grams of sulphurous acid, after long exposure to the air, are completely converted into sulphuric acid. What volume of oxygen has been absorbed from the air?

828. †How much sodium sulphite is formed if 11.16 litres of sulphur dioxide are absorbed in an excess of sodium hydrate solution?

829. What weight of sulphur dioxide is liberated when an excess of sulphuric acid is added to 7.9 grams of potassium sulphite?

830. †An excess of sulphuric acid was added to sodium sulphite when 5.58 litres of sulphur dioxide were liberated. What weight of sodium sulphite was converted into sodium sulphate?

831. An excess of barium chloride is added to 16.4 grams of sulphurous acid. What weight of barium sulphite is formed?

832. †We pass 11.16 litres of chlorine into an excess of sodium sulphite; give the names and weights of the resulting products.

SULPHUR TRIOXIDE AND SULPHURIC ACID.

833. Give the molecular weight, density, and the weight of one litre of sulphur trioxide vapour.

834. Describe fully how you would convert sulphur dioxide into sulphur trioxide.

835. Sulphur trioxide is (i.) placed on dry litmus paper, (ii.) subjected to a strong heat, (iii.) dropped into water. What is the *result in each case*?

886. Sulphurous acid is oxidised by the addition of nitric acid or hydrogen dioxide. Explain these reactions by equations.

887. The following gases are introduced by four different tubes into a large flask :—Sulphur dioxide, nitrogen dioxide, oxygen, and steam. You are required to explain the result and to illustrate the changes by equations.

888. What is the composition of the "leaden chamber crystals"? Show by an equation the action of water upon them.

889. Give the formula for Nordhausen sulphuric acid, and the method of preparing it, also show by an equation the action of heat upon it.

840. Describe any experiment you know showing the affinity which sulphuric acid has for water.

841. How would you detect the presence of sulphuric acid, or a soluble sulphate?

842. *To one test tube containing water, SO_2 has been added, to another, SO_3 has been added. How would you determine which was which?

843. Give equations showing the action of—

- (i.) Dilute sulphuric acid on zinc.
- (ii.) Strong sulphuric acid on copper.
- (iii.) Sulphuric acid on sodium chloride.
- (iv.) Sulphuric acid on potassium nitrate.

844. †What volume of oxygen is required to unite with 5.58 litres of sulphur dioxide to form sulphur trioxide, and what weight of the trioxide will be produced?

845. What is the percentage composition of sulphur trioxide?

846. What weight of sulphur trioxide may be obtained by the action of heat on 44.5 grams of Nordhausen sulphuric acid?

847. One gram of sulphuric trioxide was added to water. What weight of sulphuric acid was formed?

848. What weight of sulphuric acid may be obtained by oxidising 16 grams of sulphur dioxide in the presence of water?

849. How much sulphur will be required to produce 490 kilograms of sulphuric acid?

850. What quantity of sulphuric acid should be obtained from 100 pounds of sulphur ?

851. How many tons of sulphuric acid should be obtained from 200 tons of iron pyrites containing 48 per cent. of sulphur ?

852. If 196 tons of sulphuric acid be obtained from 200 tons of iron pyrites, what percentage of sulphur did the pyrites contain ?

853. An excess of barium chloride is added to 19.6 grams of sulphuric acid. What weight of barium sulphate is formed ?

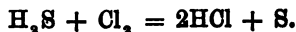
854. Some sulphuric acid was added in excess to a solution of barium chloride. The precipitate was collected on a filter, washed, dried, and weighed. It weighed 2.88 grams. What were the weights of the sulphuric acid and barium chloride used ?

855. Give the weights of caustic soda, and caustic potash, which are necessary to neutralize 4.9 grams of sulphuric acid.

856. †Give the weight and volume of the gaseous product formed when 24.5 grams of sulphuric acid are added to an excess of sodium carbonate solution.

MISCELLANEOUS.

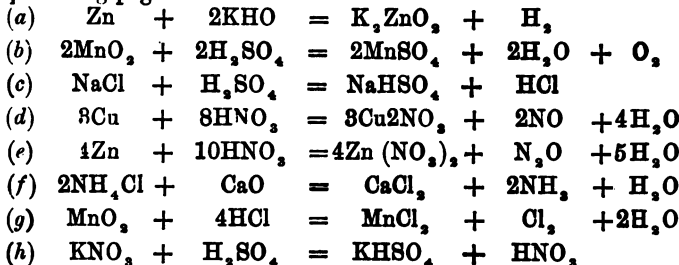
857. Explain fully the meaning of the following :—



Explanation. The following equation signifies that if sulphuretted hydrogen and chlorine be mixed together, the sulphuretted hydrogen is decomposed by the chlorine, hydrochloric acid being formed and sulphur deposited. It also illustrates the change by volume and weight, thus :—

	H_2S	+	Cl_2	=	2HCl	+	S
Name	Sulphuretted hydrogen	+	Chlorine	=	Hydrochloric acid	+	Sulphur
Volume	1 molecule	+	1 molecule	=	2 molecules	+	1 atom
Weight	$(1 \times 2) + 32$	+	35.5×2	=	$2(1 + 35.5)$	+	32
	34		71		73		

358. Explain the following equations in the manner shown on preceding page :—



359. Compare by equations the action (i.) of sulphuric acid, (ii.) of hydrochloric acid, on manganese dioxide.

360. *Into separate test glasses containing dilute hydrochloric acid I put the following substances :—Zinc, chalk, marble, common salt, charcoal, iron and gold. Mention the chemical changes which take place, and give equations.

NOTE.—Common salt, charcoal, and gold are not acted on by HCl.

361. Compare by equations the action (i.) of sulphuric acid, (ii.) of nitric acid, on copper.

362. *I add sulphuric acid to a white salt and effervescence occurs. What may this be caused by, and what tests must I apply to ascertain the nature of the gas which is evolved?

363. Compare by equations the action of sulphuric acid (i.) on common salt, (ii.) on saltpetre.

364. *Draw the graphic formulæ of the following compounds :—Ammonia, water, sulphuretted hydrogen. Give the active atomicity of each element in the following compounds :— CH_4 , SO_2 , H_2O , PF_5 , CrF_6 , and SO_2 .

365. Compare by equations the action of dilute sulphuric acid (i.) on iron, (ii.) on ferrous sulphide.

366. What is the action of water upon each of the following substances :—Hydrogen, carbonic anhydride, ammonia, sodic carbonate, chalk, and sodium?

367. *In what way would you remove a small amount of moisture from the following gases :—Air, ammonia, and carbon dioxide?

NOTE.—Gases are mostly dried by passing them through a tube containing small pieces of well dried calcium chloride, or through a tube containing small pieces of pumice stone, which have been soaked in sulphuric acid. Either of these methods would dry air, or carbon dioxide, but ammonia is absorbed largely by calcium chloride, and with sulphuric acid it unites with violence, forming ammonium sulphate. Freshly burnt lumps of lime are best for drying ammonia.

868. Find the percentage composition of the following substances :—

- (a) Potassium nitrate.
- (b) Barium sulphate.
- (c) Ammonium sulphate.
- (d) Copper nitrate.

869. How would you distinguish between potassium chlorate and potassium chloride ?

870. Define the terms "acid," "base," and "salt," and say how you would illustrate the properties of each.

871. Explain the difference between the terms "acid salt" and "neutral salt."

872. What are monobasic, dibasic, and tribasic acids ? Give examples.

873. Substances combine chemically in definite and multiple proportions. What do you understand by this statement ?

874. *Explain what is understood by Dalton's Atomic Theory.

875. By what properties is a metal distinguished from a non metal ?

876. *An excess of chlorine is passed through boiling milk of lime, and then potassic chloride added to the filtered liquid. Give the name and formula of the salt which crystallizes out.

877. *Describe accurately and fully the changes, if any, which occur when the following bodies are heated in a test tube over a gas flame :—(a) ammonium chloride (sal ammoniac), (b) sulphur, (c) graphite, (d) ammonia solution.

878. *How would you prepare the following gases CO_2 ; NH_3 ; HCl ; in a dry pure state ? Give the equations. Also, how would you fill a small bottle with each ?

879. *Name the chief properties of nitric acid. Point out in what respects nitric and sulphuric acids have similar properties, and, on the other hand, what differences there are which enable you readily to distinguish the one acid from the other.

880. *State why carbon is termed a tetrad element and oxygen a dyad element.

881. *You are required to prepare carbon monoxide and carbon dioxide from charcoal and air. Sketch the apparatus you would use, and explain how it would differ in the two cases. By what tests would you distinguish these two gases?

882. Arrange the following gases in two columns, showing which are heavier and which lighter than air:—Oxygen, nitrogen, hydrogen, ammonia, steam, nitrous oxide, and carbon monoxide.

883. Give equations showing the reactions of—

- (a) Common salt and sulphuric acid.
- (b) Copper and nitric acid.
- (c) Ammonium chloride and lime.
- (d) Potassium ferrocyanide and hot sulphuric acid.
- (e) Copper and sulphuric acid.
- (f) Manganese dioxide and hydrochloric acid.
- (g) Manganese dioxide and sulphuric acid.
- (h) Chalk and hydrochloric acid.
- (i) Oxalic acid and sulphuric acid.
- (j) Ferrous sulphide and hydrochloric acid.

884. Solutions of potassium chloride and potassium chlorate are contained in two separate test tubes: how would you distinguish between them?

885. Under what conditions, if any, do the following metals act on water:—sodium, iron, copper?

886. What is nascent hydrogen, and what is its action on ferric chloride solution?

887. Give a reason for supposing that ammonia solution contains ammonium hydroxide.

888. What is the result of passing dry ammonia gas over red hot copper oxide; also of passing air and ammonia gas together over red hot copper oxide?

889. How would you prove that carbon is present in the gas which is given off on the addition of HCl to washing soda (carbonate of soda)?

390. Compare the heating powers of a jet of hydrogen burning in air and oxygen. Explain the reason for the difference.

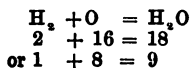
"The temperature of a hydrogen flame burning in air has been estimated at about 2080°C, but when the flame is fed with pure oxygen its temperature rises to upwards of 4000°C. This is easily explained by the fact that in atmospheric air the oxygen is mixed with four times its bulk of nitrogen, which contributes nothing to the chemical action, and which, being raised to the same temperature as the other gases present, consumes a great deal of heat."—*Tilden's Chemical Philosophy*, p. 100.

391. Explain what is meant by heat of combustion, and define the term "heat unit."

"It has been shown by numerous experiments that when the same weight of the same substance burns to form the same products of combustion, a constant amount of heat is invariably evolved whether the combustion takes place slowly or quickly."—*Roscoe and Schorlemmer's Treatise on Chemistry*, vol. 1, p. 189.

This constant amount of heat is termed the "heat of combustion."

Thus, when 1 part by weight of hydrogen combines with 8 parts by weight of oxygen to form 9 parts by weight of water,



the quantity of heat which is liberated is sufficient to raise 84,462 parts by weight of water from 0° to 1° C., no matter how slowly or quickly the combustion proceeds.

Hence, 84,462 heat units is said to be the calorific power of hydrogen, and this power for different substances is measured by the amount of water which can be raised from 0° to 1° C by the combustion of a unit weight of the substance.

Calorific power of Hydrogen burnt in oxygen	=	84,462.
Do. do. Carbon do. do.	=	8,080.
Do. do. Carbon monoxide do. do.	=	2,403.

Definition of "heat unit."—A "heat unit" is the amount of heat required to raise a unit weight of water from 0° to 1° C.

392. What weight of water would be heated from 0° to 1° C by the combustion of 10 grams of hydrogen in oxygen?

393. What weight of water would be heated from 0° to 10° C by the combustion of 1 gram of hydrogen in oxygen?

394. How many grams of hydrogen would be required to burn in oxygen so as to raise 17,231 grams of water from 0° to 5° C?

895. What weight of water would be raised from 0° to 2°C by the combustion of 10 grams of carbon in oxygen?

1 gram of carbon burnt in O will raise 8080 grams of water from 0° to 1°C .
 1 gram of carbon burnt in O will raise 4040 grams of water from 0° to 2°C .
 10 grams of carbon burnt in O will raise 40400 grams of water from 0° to 2°C .

896. Calculate the weight of water raised from 0° to 3°C by the combustion of 10 grams of carbon monoxide in oxygen.

Find the empirical formulæ of the following substances from their percentage composition :—

897. Hydrogen - - 5.88
 Oxygen - - 94.12

 100.00

898. Manganese - - 72.05
 Oxygen - - 27.95

 100.00

899. Calcium - - 40.00
 Carbon - - 12.00
 Oxygen - - 48.00

 100.00

400. Sodium - - 39.82
 Chlorine - - 60.68

 100.00

401. Iron - - 46.67
 Sulphur - - 53.33

 100.00

402. Copper - - 25.45
 Sulphur - - 12.82
 Oxygen - - 25.66
 Water - - 36.07

 100.00

403. Potassium - - 88.61
 Nitrogen - - 11.86
 Oxygen - - 47.53

 100.00

404. Hydrogen - - 2.04
 Sulphur - - 82.65
 Oxygen - - 65.31

 100.00

405. Phosphorus - - 18.90
 Sodium - - 42.07
 Oxygen - - 39.03

 100.00

APPENDIX.

METRIC SYSTEM OF WEIGHTS AND MEASURES.

MEASURES OF LENGTH.

Kilometre	=	1000 metres	=	39370	inches.
Hectometre	=	100 metres	=	3937·0	"
Decametre	=	10 metres	=	393·70	"
Metre	=		=	39·370	"
Decimetre	=	$\frac{1}{10}$ metre	=	3·9370	"
Centimetre	=	$\frac{1}{100}$ metre	=	·39370	"
Millimetre	=	$\frac{1}{1000}$ metre	=	·039370	"

NOTE.—The abbreviation c.c. = cubic centimetre.

MEASURES OF CAPACITY OR VOLUME.

Kilolitre	=	1000 litres	=	1760·7	pints.
Hectolitre	=	100 litres	=	176·07	"
Decalitre	=	10 litres	=	17·607	"
Litre	=		=	1·7607	"
Decilitre	=	$\frac{1}{10}$ litre	=	0·17607	"
Centilitre	=	$\frac{1}{100}$ litre	=	0·017607	"
Millilitre	=	$\frac{1}{1000}$ litre	=	0·0017607	"

NOTE.—1 litre = 1 cubic decimetre.

1 millilitre = 1 c.c. (cubic centimetre).

MEASURES OF WEIGHT.

Kilogram	=	1000 grams	=	15432·3	grains.
Hectogram	=	100 grams	=	1543·23	"
Decagram	=	10 grams	=	154·323	"
Gram	=		=	15·4323	"
Decigram	=	$\frac{1}{10}$ gram	=	1·54323	"
Centigram	=	$\frac{1}{100}$ gram	=	0·154323	"
Milligram	=	$\frac{1}{1000}$ gram	=	0·0154323	"

NOTE.—An ounce avoirdupois equals 35·235 grams.

ANSWERS.

24. 1844 grams, 896 grams, 4928 grams.
25. 11.16 litres, 5.58 litres, 2.79 litres, 27.0 litres, 111.6 litres.
26. 90 grams of water.
27. 28 grams of sodium.
28. 224 grams of caustic potash.
29. 4 grams of hydrogen.
30. 6 grams of hydrogen ; 408 grams of zinc chloride.
31. 89.28 litres of hydrogen.
32. 5.82 grams of zinc.
33. 267.84 litres of hydrogen.
34. 696 grams of magnetic oxide of iron.
35. 56 grams of caustic potash.
41. 2.15 grams, 14.88 grams, 7.88 grams.
42. 6975 litres, 11.16 litres, 16.74 litres, 69.75 litres.
43. 600 grams of mercury.
44. 96 grams of oxygen.
45. 867.5 grams of potassium chlorate.
46. 192 grams of oxygen.
47. 12.416 grams of potassium chloride.
48. 122.5 grams of potassium chlorate.
49. 188.92 litres of oxygen.
50. 86 grams of water.
51. 112 grams of oxygen.

52. 64 grams of oxygen.
53. (i.) 7.407; (ii.) 88.88; (iii.) 65.306.
54. 897.5 grams of copper oxide.
74. 124 litres of H; 62 litres of O.
75. 18 grams of hydrogen.
76. 558 litres of oxygen.
77. 11.25 grams of water.
78. 180 grams of water.
79. 127 grams of copper.
80. 28 c.c. of steam.
81. 30 vols., viz :—26 vols. of steam and 4 vols. of hydrogen.
82. $27\frac{1}{2}$ c.c., viz. :—25 c.c. of steam and $2\frac{1}{2}$ c.c. of oxygen.
83. 5 litres of oxygen; 8.064 grams of water.
84. 864 grams of mercuric oxide.
93. 68 grams of hydrogen peroxide.
94. 507 grams of barium dioxide.
95. 146 grams of HCl; 68 grams of H_2O_2 .
96. 144 grams of oxygen.
97. 94.117.
105. 4.77 grams; 7.95 grams; 81.8 grams.
106. 11.16 litres, 22.32 litres, 89.28 litres.
107. 48.5 grams of manganese dioxide.
108. 78 grams of hydrochloric acid; 63 grams of manganous chloride.
109. 180.5 grams of manganese dioxide.
110. 35.5 grams; 11.16 litres.
111. 5.24 grams.
118. 3.27 grams; 5.72 grams; 8.99 grams.
119. 11.16 litres; 1.22 litres; 8.05 litres.
120. 117 grams of sodium chloride.

121. 49 grams of sulphuric acid ; 60 grams of acid sodium sulphate.
122. By weight, 2 grams of H, 71 of Cl ; by volume, 22·82 litres of each.
123. 284 grams of NaCl.
126. 8 grams.
135. 4·928 grams, 10·841 grams, 19·712 grams.
136. 22·82 litres, 507 litres, 2·788 litres.
137. 48 grams, or 88·48 litres of oxygen.
138. 200 litres of air.
139. 44 grams of CO_2 , 16 grams of oxygen remain.
140. 22·82 litres of CO_2 , 22·82 litres of oxygen required.
141. 11·16 litres of CO_2 , 84 litres of oxygen remain.
142. 56 grams of lime.
143. 219 grams of HCl, 800 grams of marble.
144. 4·48 grams.
145. 86 grams of water.
146. Calcium carbonate, difference 2·5 grams.
147. 100 grams of HCl, difference 15·876 grams of CO_2 .
148. 2·2 grams of carbon dioxide.
149. 106 grams of sodium carbonate.
157. 8·186 grams, 6·899 grams, 9·408 grams.
158. 22·82 litres, 5·58 litres, 89·857 litres.
159. 252 grams of carbon monoxide, 18 grams of hydrogen.
160. 42·85 per cent in CO ; 27·27 per cent. in CO_2 .
161. 11·16 litres, or 14 grams of carbon monoxide.
162. 1 litre, or 1·97 grams.
163. 4 volumes of carbon dioxide ; 2 volumes of oxygen remain uncombined.
164. 66 grams of carbon monoxide ; 8 grams of carbon remain.

165. 45 grams of oxalic acid.
166. 56 grams or 44·64 litres of CO ; 88 grams or 44·64 litres of CO₂.
167. 42 grams, or 38·48 litres of CO.
168. 2·19 grams of K₄FeC₆N₆.
169. ·7168 grams, 8·942 grams, 6·092 grams.
170. 11·16 litres, 1·895 litres, 7·6725 litres.
171. 128 grams, or 89·28 litres of oxygen.
172. 38 grams of carbon dioxide, 27 grams of oxygen.
178. 100 litres of air.
174. 84 litres of oxygen, 400 litres of air.
179. 8·186 grams, 6·899 grams, 10·66 grams.
180. ·797 litres, 11·16 litres, 16·74 litres.
181. 2·86 grams of ammonium nitrite.
182. 44·64 litres of nitrogen.
188. 6·46 grams, 10·98 grams, 12·92 grams.
189. 22·82 litres, 5·58 litres, 77·89 litres.
190. 237 vols. of nitrogen.
191. 48·26 grams.
192. The flask containing the equal volumes of N and O ; difference, 52 milligrams.
198. 8·625 grams of carbon.
194. 17·82 grams of phosphorus.
195. 46 grams of sulphur dioxide.
196. 2·875 grams.
197. 42 litres of hydrogen.
198. 100 litres of air.
199. 114·28 grams of copper oxide ; 77 grams of nitrogen.
200. 484·78 kilograms of air.
209. 1·14 grams, 4·18 grams, 7·61 grams.

210. 5.58 litres, 22.82 litres, 82.82 litres.
211. 17 grams.
212. Nitrogen 82.85
 Hydrogen 17.65
100.00
218. $\frac{1}{2}$ litre of N, $1\frac{1}{2}$ litres of H ; $\frac{1}{2}$ litres of O.
214. Water, and 98 c.c. of N.
215. 89.28 litres.
216. 2.896 grams of NH_4Cl .
217. 56 grams of nitrogen, 108 grams of water.
218. 68 grams of ammonia.
219. 44.64 litres of ammonia, 132 grams of ammonium sulphate.
220. Hydrochloric acid requires the larger quantity. The difference is 19.59 grams.
228. 7.056 grams, 15.528 grams, 28.224 grams.
229. 11.16 litres, 22.82 litres, .854 litres.
230. $15\frac{1}{2}$ pounds.
231. 24.5 grams.
232. 28.67 grams of potassium.
233. 20 grams of caustic soda ; 42.5 grams of sodium nitrate.
240. 4.928 grams, 14.784 grams, 19.712 grams.
241. 11.16 litres, 16.74 litres, .507 litres.
242. 1.52 grams.
243. 7 litres of nitrogen.
244. The 10 litres are increased to 15 litres.
245. 8.584 grams of oxygen.
246. 11 grams of CO_2 ; 22 grams of nitrous oxide.
247. 8 grams of SO_2 ; 5.58 litres of free N.
248. 3 grams of carbon.
249. 220 grams of nitrous oxide.

250. 11·16 litres of H, and 9 grams of water.
251. 4·5 grams of water.
252. 14 grams, or 11·16 litres of N.
253. 5·5 grams of nitrous oxide.
254. 8·584 grams of ammonium nitrate.
255. 11 grams, or 5·58 litres of N₂O.
256. 890 grams of zinc.
257. 28·686 grams of nitric acid.
264. 15 grams, 8·75 grams, 1·844 grams.
265. 22·82 litres, 5·58 litres, 74·4 litres.
266. 80 grams.
267. 58·88 per cent.
268. 62 grams of phosphorus.
269. 15 grams of nitric oxide, 5·58 litres of nitrogen.
270. 111·6 litres of nitric oxide.
271. 126 grams of nitric acid.
272. 95·25 grams of copper, 281·25 grams of copper nitrate.
273. ·8149 grams of nitric oxide.
290. 8·808 grams, 8·877 grams, 15·282 grams.
291. 22·82 litres, 5·58 litres, ·656 litres.
292. 48·26 grams.
293. 6 grams, or 66·96 litres of hydrogen.
294. Iron 68·64
 Sulphur 86·86
 100·00
295. 8·5 grams, or 5·58 litres of H₂S.
296. 11 grams of ferrous sulphide.
297. 294 grams of sulphuric acid.
298. 18·25 grams of hydrochloric acid.
299. 16·74 litres of oxygen.

300. 4.25 grams of H_2S .
301. 16 grams of SO_2 , 4.5 grams of H_2O .
302. 3.82 grams of copper sulphide.
303. 27.2 grams of zinc chloride.
304. 6.8 grams of sulphuretted hydrogen.
305. 17.75 grams of chlorine ; sulphur, 8 grams ; hydrochloric acid, 18.25 grams.
316. 32 grams, 2.867 grams, 16 grams.
317. .523 litres, 2.79 litres, 3.66 litres.
318. Air, 7.21 grams ; carbon dioxide, 11 grams.
319. 1 gram of oxygen ; .697 litres of sulphur dioxide.
320. 15.875 grams of copper.
321. 24.5 grams of sulphuric acid.
322. 2.56 of sulphur dioxide.
323. 6.4 grams of sulphur dioxide ; 29.6 grams of mercuric sulphate.
324. 64 grams, or 22.82 litres of SO_2 ; 22 grams, or 11.16 litres of CO_2 .
325. .9 grams of water.
326. 123 grams of sulphurous acid.
327. 5.58 litres of oxygen.
328. 63 grams of sodium sulphite.
329. 3.2 grams of sulphur dioxide.
330. 31.5 grams of sodium sulphite.
331. 43.4 grams of barium sulphite.
332. Sodium sulphate, 71 grams ; sodium chloride, 58.5 grams ; sulphur dioxide, 32 grams.
344. 2.79 litres of oxygen ; 20 grams of sulphur trioxide.
345. Sulphur, 40 ; oxygen, 60.
346. 20 grams of sulphur trioxide.
347. 1.225 grams of sulphuric acid.

848. 24.5 grams of sulphuric acid.
849. 160 kilograms of sulphur.
850. 806.25 pounds of sulphuric acid.
851. 294 tons of sulphuric acid.
852. 82 per cent of sulphur.
853. 46.6 grams of barium sulphate.
854. .98 grams of sulphuric acid; 2.08 grams of barium chloride.
855. 4 grams of caustic soda, 5.6 grams of caustic potash.
856. 11 grams, or 5.58 litres of carbon dioxide.
- 868.
- | | | |
|-----|------------|--------------|
| (a) | Potassium | 38.61 |
| | Nitrogen | 13.86 |
| | Oxygen ... | 47.53 |
| | | <hr/> 100.00 |
- | | | |
|-----|------------|--------------|
| (b) | Barium ... | 58.80 |
| | Sulphur... | 13.78 |
| | Oxygen ... | 27.47 |
| | | <hr/> 100.00 |
- | | | |
|-----|------------|--------------|
| (c) | Nitrogen | 21.21 |
| | Hydrogen | 6.06 |
| | Sulphur | 24.24 |
| | Oxygen ... | 48.49 |
| | | <hr/> 100.00 |
- | | | |
|-----|------------|--------------|
| (d) | Copper ... | 38.87 |
| | Nitrogen | 14.98 |
| | Oxygen... | 51.20 |
| | | <hr/> 100.00 |
892. 844,620 grams of water.
893. 8446.2 grams of water.
894. 2.5 grams of hydrogen.
895. 40400 grams of water.

396. 8010 grams of water.

397. H_2O_2 .

398. Mn_2O_3 .

399. CaCO_3 .

400. NaCl .

401. FeS_2 .

402. $\text{CuSO}_4 + 5 \text{H}_2\text{O}$.

403. KNO_3 .

404. H_2SO_4 .

405. Na_3PO_4 .

OUTLINES OF EXPERIMENTS

SUITABLE FOR ILLUSTRATING

Elementary Instruction in Chemistry.

PREPARED BY

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BY DIRECTION OF THE LORDS OF THE COMMITTEE OF COUNCIL ON EDUCATION
FOR THE ASSISTANCE OF TEACHERS OF SCIENCE SCHOOLS AND CLASSES.

SUBJECT X. ELEMENTARY STAGE.

The following notes have been prepared as some guide to the teachers as to the general character of the course of instruction expected in the elementary stage. It includes instruction that should on no account be omitted, but must be considered rather as suggestive than exhaustive.

I.—COMBUSTION AND CHEMICAL COMBINATION.

1. Burn a taper in a clean glass bottle. Show the presence of a colourless gas, differing in properties from common air by yielding a turbidity with lime water.

2. Hold a bright glass over a burning candle and show the formation of water. Explain what is meant by chemical change, and state that chemistry is an experimental science.

3. Make similar experiments with a petroleum or paraffin lamp.

4. Show that coal gas also yields the same products by passing the products of combustion through lime water and by collecting the water.

5. Explain the difference between mechanical mixture and chemical combination; and illustrate by a mixture of finely divided copper and flour of sulphur, and the effect of heat upon the same.

6. Experiment to show that chemical change consists of a change in the properties of matter and that no loss of matter takes place. Suspend lamp-chimney, partly filled with lumps of caustic soda, from the arm of a balance. Place short piece of candle in the lower part of the glass and counterbalance. Light the candle. Explain the increase in weight.

7. Heat is evolved when chemical combination takes place. Pour water on to quicklime. Refer also to experiments 1 and 3.

8. Combustibles and supporters of combustion. The purely relative character of these terms. Ordinary combustion the union of atmospheric oxygen with a body termed the combustible, or with one or more of its constituents, heat being developed, as in all cases where two or more bodies combine. Illustrate by *showing that air will burn in coal gas just as well as coal gas will burn in air.*

II.—AIR.

1. Existence of atmosphere, felt in winds.
2. Weight of air shown by means of a flask exhausted by the air pump.
3. Burn phosphorus in air.
4. Burn phosphorus in confined volume of air and show diminution in bulk.
5. Show that some diminution takes place slowly when a stick of phosphorus is exposed to air at ordinary temperatures.
6. Test residual gas (N) with a burning taper.
7. Show that phosphorus burnt in air increases in weight.
8. Allow iron borings moistened with sal ammoniac to rust in a confined volume of air and introduce burning taper into residual gas (N).
9. Show that iron filings, suspended by a magnet hanging on one scale of a balance, increase in weight on heating.
10. Strongly heat the red substance which may be formed by gently heating mercury in the air. Collect and test the gas (O) with a glowing splinter of wood.
11. Add the gas thus obtained to the residue obtained in experiment 4 or 8 so as to make up the original volume of air, and show that a taper burns in this mixture as in common air.
12. Refer to numbers giving exact analysis of air, calling especial attention to the fact that it varies slightly in composition.

Also explain that no obvious change, such as increase of temperature or alteration of bulk occurs when oxygen and nitrogen are mixed. Also that air has the properties of a mixture, and that when water is shaken up with air a portion of that air dissolves, the residue being found to contain relatively less oxygen than the original air, whilst the dissolved portion contains relatively more oxygen, and that this could not be the case if the air were a compound. Consequently it is a mixture and not a chemical compound.

13. It is important that these experiments should be made and their explanation given so as to teach the student how the composition of air is ascertained by experiment, and in a similar manner how oxygen was discovered by Priestley, and how the composition of the air and the part which oxygen plays in the phenomena of combustion were experimentally demonstrated by Lavoisier.

III.—EFFECTS OF ANIMAL AND VEGETABLE LIFE UPON THE ATMOSPHERE.

1. Show that by drawing air into the lungs through lime-water a very faint, if any, precipitate is produced; but that on expiring air from the lungs through another portion of lime-water a copious precipitate is soon formed.
2. Show the production of carbon dioxide by the oxidation of ordinary articles of food, as by heating small quantities of the dried substance, such as sugar, bread, or meat, with copper oxide.
3. Show that carbon dioxide exists in the air by pouring clear lime-water into a shallow vessel exposed to air, and explain that this small quantity of carbon dioxide serves as the main food of the plants that grow on the earth.
4. Expel air from water by boiling, and explain how fish and aquatic plants are thus provided with oxygen and carbonic acid.
5. Explain that plants eliminate and animals require oxygen. That animals take in oxygen from the air, and give out carbonic acid. That plants possess the power under the influence of light of assimilating carbon from carbon dioxide and liberating the oxygen. Explain that thus the balance of oxygen and carbon dioxide in the atmosphere is maintained.
6. Illustrate the action of plants by the formation of bubbles of oxygen when a fresh plant is exposed to the action of light in water containing carbonic acid in solution.

IV.—WATER.

1. Illustrate the three states of matter, the solid, the liquid, and the gaseous, with ice, water, and steam, and point out that the difference is caused by increase or diminution in the amount of heat present.
2. Composition of water. Decompose water by the electric current. Collect the two gases separately in a voltameter, and exhibit their properties.
3. Show formation of water by explosion of a mixture of hydrogen (two volumes) and oxygen (one volume) in a soda-water bottle.
4. Explode soap bubbles inflated by a mixture of hydrogen and oxygen in the above proportions.
5. Throw potassium or sodium into water, and collect the hydrogen.
6. Pass steam over red-hot iron, collect the gas and show that it is hydrogen.
7. Show that the same gas may be obtained by dissolving zinc clippings or iron turnings in dilute sulphuric acid.
8. Demonstrate the properties of hydrogen :
 - (a.) Its combustibility.
 - (b.) Its lightness.
 - (c.) That a candle will not burn in it.
 - (d.) That water is formed when it burns in air.
9. Composition of water. Pass oxygen over red-hot copper, and show by weighing before and after that the weight increases.
10. Pass hydrogen over the copper oxide thus produced, heating gently. Collect the water, and show that the copper oxide has been entirely reduced, the tube weighing the same as before passing the oxygen through it.
11. Determine the composition of water by weight by passing dry hydrogen over half an ounce of copper oxide, and collecting the water in a weighed chloride of calcium tube. Show approximately that water contains two parts by weight of hydrogen to 16 parts by weight of oxygen.
12. Note the first law of chemical combination that chemical compounds, such as water, always contain their components in the same unvarying proportions.
13. Contrast the properties of water with those of its constituents on the one hand, and the properties of air, with those of its constituents on the other.
14. Call attention to air and water as illustrations of the difference between a mixture and a compound, and quote oxygen, nitrogen, and hydrogen as examples of elementary bodies.
15. Separation of impurities from water by filtration and distillation. Preparation of fresh water from salt water.
16. Experiments illustrating solution and crystallization. Soluble substances, as sugar, washing soda, alum; slightly soluble substances, as gypsum or plaster of Paris; insoluble substances as chalk, flint, and sand.
17. Crystallize carbonate of soda, and sulphate of copper

V.—OXYGEN AND OZONE.

1. Prepare oxygen by heating—
 - (a.) Oxide of mercury.
 - (b.) Potassium chlorate.
 - (c.) Mixture of potassium chlorate, and either manganese dioxide, copper oxide or ferric oxide.
2. Show the re-ignition of a splinter of red-hot wood and glowing wick of taper.
3. Burn charcoal in oxygen, and show the formation of carbon dioxide.
4. Burn phosphorus simultaneously in air and in oxygen.
5. Burn watch spring in oxygen.

6. Show that iron does not rust in dry oxygen.
7. *Ozone*.—Describe and demonstrate the formation of ozone by submitting oxygen to the silent electric discharge.
8. Describe and demonstrate the properties of ozone which distinguish it from ordinary oxygen, such as its action on metallic mercury, on indigo solution, or on potassium iodide and starch. Also its change to ordinary oxygen when passed through a heated glass tube.
9. Explain the difference in density between oxygen and ozone.

VI.—COMBINING WEIGHTS; NAMES AND SYMBOLS OF THE ELEMENTS;
CHEMICAL CALCULATIONS, &c.

1. Exhibit list of the elements, distinguishing (by means of the type) the non-metals from the metals; and, again, the more commonly occurring metals from those which are rarer.
2. Describe the occurrence of these elements in the air, in the sea, and in the solid crust of the earth.
3. Write up the results of the quantitative analysis of potassium chlorate. Explain that this is the result of experiment, and demonstrate the fact that, when heated, an unalterable weight of oxygen is given off and a given unalterable weight of potassium chloride remains behind.
4. Dissolve a crystal of pure chlorate of potassium, and the residue of chloride from heating chlorate, in two separate glasses, and show the difference in the reaction with silver chloride.
5. Explain the meaning of the term chemical symbol, and chemical formula of a salt, referring afterwards to the combining weights of the elements.
6. Explain the mode of determining the formula from the percentage composition.
7. Method of calculating the quantity of oxygen from potassium chlorate (and from manganese dioxide).

VII.—ACIDS, BASES, AND SALTS.

1. Burn sodium in oxygen; dissolve the product in water; give the formula of the oxide. Express the action of water upon it by an equation.
2. Act on water with sodium, and collect the hydrogen. Explain by equation that the same substance, sodium hydrate or hydroxide, or caustic soda, is formed, as in Experiment 1.
3. Burn sulphur in a current of oxygen and show that the product fumes slightly in the air. Explain that it is a mixture of sulphur dioxide and trioxide. Pass the gas thus obtained into water.
4. Add litmus solution to the solutions thus obtained in Experiments 2 and 3, and show that on adding the one solution to the other the colour is changed, or a point is reached where a further addition of the one has no effect, whereas a minute addition of the other at once changes the colour. Explain the action by an equation.
5. Explain that the compound formed from the sodium oxide and water is termed an alkali, or alkaline, or basic hydroxide, and the oxide from which it is formed an alkaline or basic oxide; that the compound formed from the sulphur dioxide and water is termed an acid hydroxide or acid, and the original oxide an acid forming oxide or anhydride.
6. Explain that sodium hydroxide and sulphurous acid may be taken as representative of the two classes into which hydroxides are divided.
7. Explain that by the action of the one upon the other a salt is formed. Exhibit a white crystalline salt, *e.g.* sodium sulphate.

VIII.—HYDROGEN.

1. Prepare hydrogen by the action of dilute sulphuric acid on zinc.
2. Show that it is not obtained by the use of pure zinc (amalgamated zinc is best used), and illustrate the effect of impurity by adding a drop or two of a lead or copper salt.
3. Prepare hydrogen by dissolving zinc or aluminium in sodium hydroxide.
4. Point out that whereas sodium displaces hydrogen from water at ordinary temperature, and that iron does so at a red heat, copper is without any action at any temperature.
5. Give equations for the various methods here indicated for obtaining hydrogen.
6. Explain fully in detail the methods of chemical calculation and make the pupils thoroughly understand the methods of calculating quantities.
7. Demonstrate the physical properties of hydrogen, especially its lightness and diffusibility.
8. Compare the heating powers of jet of hydrogen burning in air and in oxygen. Explain the difference.
9. Describe and (if possible) demonstrate the construction and use of the oxhydrogen blow-pipe.
10. Explain what is meant by heat of combustion, and define the term "heat-unit." Show for this purpose side by side 18 grams of water and the quantity of water which would be raised 1° C. in temperature by the heat developed in the formation of this quantity of water from its elements.
11. Point out that hydrogen is a powerful reducing agent, illustrating this by the reduction of oxide of iron.
12. Show that nascent hydrogen, or hydrogen at the moment of its liberation from its compounds, frequently produces effects that hydrogen in the free state does not. Bubble hydrogen through ferric chloride solution and show that no discolourization takes place. Place it in contact with zinc and dilute sulphuric acid and the colour disappears.
13. Explain the term nascent, as applied to hydrogen and other gas at the moment of its liberation from one of its compounds, and distinguish between the atom of nascent hydrogen and the molecule of free hydrogen.

IX.—HYDROCHLORIC ACID AND CHLORINE.

1. Explain with equation and show the action of sulphuric acid on common salt. Collect the escaping gas by downward displacement and show its solubility in water.
2. Hold piece of paper dipped in ammonia solution in the gas.
3. Saturate water with the gas, noting that its volume increases and that considerable heat is developed.
4. Exhibit the effects produced by adding the solution to litmus and to silver nitrate solution.
5. Show that it has no action on indigo, or on a mixture of potassium iodide and starch solution.
6. Pass the gas over red-hot iron and show the production of hydrogen.
7. *Chlorine*.—Heat oxide of manganese with the solution of hydrochloric acid obtained in Experiment 3, and collect several jars of the escaping chlorine by downward displacement. Give the equation.
8. Pass some of the gas into water. Exhibit the yellow colour of the solution and show that it precipitates silver nitrate and bleaches litmus and indigo.
9. Burn a jet of hydrogen in chlorine. Show the disappearance of the yellow coloured gas.

10. Moisten some paper with a few drops of turpentine and throw it into a jar of chlorine. Point out the formation of hydrochloric acid and the deposition of carbon.

11. Explode a mixture of equal volumes of hydrogen and chlorine.

12. Point out how these experiments show that the gas produced in Experiment 1 is a compound of chlorine and hydrogen. Give the symbol and atomic weight of chlorine, and state the composition of hydrochloric acid gas by weight and volume.

13. Explain the production of chlorine from common salt, sulphuric acid, and manganese dioxide. Give equations, and instruct the students in the calculations of quantities.

14. Show the combustion in chlorine, of phosphorus, antimony, and copper, and demonstrate its power to displace bromine and iodine from their compounds with metals.

15. Electrolysis of hydrochloric acid solution and explain the fact of the evolution of equal volumes of its constituent gases.

16. Explain the bleaching action of chlorine as being due to the readiness with which it combines with hydrogen and that it thus acts as an oxidizing agent. In illustration of this, show that a piece of dry turkey red cloth when placed in dry chlorine is not bleached.

X.—NITROGEN AND AMMONIA.

1. The production of nitrogen from the air and the examination of its properties may here be repeated.

2. Describe and (if possible) demonstrate the production of ammonia by passing sparks from an induction coil or electric machine through a mixture of ammonia and hydrogen. Explain that the reaction is not complete unless the ammonia is withdrawn as it is formed, owing to the fact that ammonia is readily decomposed by heat.

3. Prepare ammonia by heating an ammoniacal salt with slaked lime. Collect by upward displacement and over mercury, and show extreme solubility in water.

4. Demonstrate and explain its combination with hydrochloric acid, and show the volatility of sal ammoniac.

5. Show that the aqueous solution of ammonia behaves in the same way as a solution of sodium hydroxide, turning red litmus blue, neutralising acids, and forming precipitates in solutions of metals (copper, iron and zinc salts, for example) of the same composition as those produced by sodium hydroxide. Explain that on this account it is considered that the ammonia solution contains ammonium hydroxide.

6. Pass dry ammonia gas over red hot copper oxide and show the production of water and metallic copper.

7. Pass air and ammonia gas simultaneously over red hot copper as a method of preparing nitrogen.

XI.—NITRIC ACID AND THE OXIDES OF NITROGEN.

1. Explain on the blackboard the composition by weight of the five distinct oxides of nitrogen as illustrative of the law of chemical combination in multiple proportions, and as a deduction from this, explain Dalton's atomic theory and state clearly what is meant by an atom. Demonstrate with a series of blocks labelled with the symbols of the different elements how this explains the observed facts of combination in multiple proportions.

2. Make clear to the student the difference between atom and molecule, and explain atomic weight and molecular weight of (1) hydrogen; (2) oxygen; ozone; (3) chlorine; and then of compounds such as (4) hydrochloric acid; (5) water; (6) ammonia.

3. Describe and (if possible) demonstrate the formation of the red fumes or nitric peroxide on passing an electric spark through air.

4. Preparation of nitric acid from nitre and sulphuric acid. Explain the reaction by an equation.

5. Calculation of quantities to be carefully gone into.

6. Exhibit nitre, nitrate of soda, sulphate and bisulphate of potash, and sulphate and bisulphate of soda.

7. Show the oxidizing action of nitric acid by dropping it on to some red hot charcoal.

8. Oxidizing action of nitric acid on metallic tin and metallic copper.

8A. Deflagrate mixture of nitre and charcoal.

9. Show the decomposition of nitric acid when heated by dropping it into the bowl of a clay tobacco pipe, the stem of which is strongly heated, collecting the gas over water and testing with a flaming splinter of wood.

10. Heat potassium nitrate and collect the gas (O).

11. Prepare nitric oxide from residue in experiment 10 by treating with dilute sulphuric acid. Explain decomposition of nitrous acid into nitric oxide and nitric acid.

12. Prepare nitric oxide by action of nitric acid on copper turnings. Collect the gas. Explain the reaction.

13. Exhibit the direct combination of nitric oxide with oxygen. Note the formation of red fumes of nitrogen peroxide and their immediate absorption by water.

14. Show that flame of a taper is extinguished in nitric oxide, and that feebly burning phosphorus is also extinguished, but that brightly burning phosphorus continues to burn, and with greater brilliance than in ordinary air. Explain this.

15. Preparation of nitrous oxide. Neutralise nitric acid with ammonia. Evaporate the solution and obtain the solid salt. Show the preparation of nitrous oxide with this residue. Collect the gas over warm water. Give equation. Explain that nitrous oxide is readily soluble in cold water.

16. Show that like oxygen, nitrous oxide supports the combustion of a taper, and explain that this is caused by the decomposition of the gas, and the union of the constituents of the taper with the oxygen of the nitrous oxide, and liberation of the nitrogen.

17. Also show that phosphorus and strongly ignited sulphur burn in the gas, but that feebly ignited sulphur is extinguished. Explain this.

18. Point out the distinction between nitrous oxide and oxygen: (1) the solubility of nitrous oxide in cold water, (2) the production of nitrogen when bodies burn in it, (3) the fact that nitric oxide does not produce with it red fumes, as is the case with oxygen.

19. Prepare nitrogen from ammonium nitrite (*i.e.* a mixture of potassium nitrite and ammonium chloride).

20. Explain how, in the above experiments, the gradual deoxidation of nitric acid yields the several oxides of nitrogen, and lastly, nitrogen itself.

XII.—SULPHUR.

1. Exhibit the different forms of sulphur; flour of sulphur, brimstone or stick sulphur, and crystallized native sulphur.

2. Dissolve sulphur in bisulphide of carbon, and obtain crystals by spontaneous evaporation. Indicate the identity of this form with the naturally occurring crystals, and its difference from that obtained by fusing sulphur and allowing the mass to cool.

3. Explain what is meant by allotropic modification, and point out how the one form of crystal passes into the other.

4. Show the effect of heat upon sulphur melted in a flask. Contrast the brittle mass derived from cooling the sulphur after heating slightly above its melting point by pouring into cold water, with the plastic mass obtained when cooled in the same way from a high temperature. Point out the changes which occur as the temperature rises, and exhibit the red vapour of sulphur.

5. Show combustion in sulphur vapour. Insert a coil of copper wire into the sulphur vapour, and show that combination occurs.

6. Distil sulphur in a small retort.

7. Pass hydrogen through boiling sulphur, and demonstrate the formation of hydrogen sulphide by its blackening action on lead paper.

8. Exhibit ferrous sulphide and galena (lead sulphide). Prepare hydrogen sulphide (sulphuretted hydrogen) from the former by the action of dilute sulphuric acid. Collect by displacement and prepare a solution of the gas in water.

9. Show the combustible nature of hydrogen sulphide by burning a jar of the gas, and point out the deposition of sulphur due to incomplete combustion. Demonstrate and explain the decomposition of hydrogen sulphide by chlorine, and show the deposition of sulphur when its solution is allowed to stand exposed to the air and light.

10. Demonstrate the value of hydrogen sulphide as a means of separating the metals into groups, by adding the solution or passing the gas into solutions of the various metals, as for example, arsenious acid, copper sulphate, lead nitrate, antimony chloride, zinc sulphate, ferrous sulphate, and magnesium sulphate.

Write down the equations in each case.

11. Prepare sulphur dioxide by heating copper with sulphuric acid and collect the gas.

12. Illustrate the condensation of a gas into a liquid by passing sulphur dioxide into a glass tube surrounded by a freezing mixture of ice and salt.

13. Pass the gas into water and demonstrate the acid properties of the solution.

14. Prepare sulphur trioxide from fuming Nordhausen sulphuric acid. Add it to water and compare its behaviour with that of the dioxide under similar circumstances.

15. Describe the formation in the above experiment of sulphuric acid, explain the properties of oil of vitriol, demonstrating its affinity for water as exhibited by the great heat evolved when the two liquids are mixed.

16. Explain the barium chloride test for sulphuric acid.

17. Add barium chloride to a solution of sulphurous acid, and then nitric acid.

18. Explain that in consequence of the readiness with which sulphurous acid takes up oxygen it acts as a bleaching agent and as a powerful reducing agent.

XIII.—CARBON.

1. Show the presence of carbon (charcoal) in wood by carbonizing a splinter of wood in a test tube; and in white sugar, by pouring strong sulphuric acid on to a syrupy solution.

2. Describe the properties and modes of occurrence of the three allotropic modifications of carbon: (a) the amorphous form (lampblack and charcoal), and the two crystalline forms, (b) graphite, and (c) diamond. Describe the octohedral forms of the crystal of diamond and show glass or wood models.

3. Explain that the same weight of each of these substances when burnt gives the same weight of the same product (carbon dioxide).

4. Calculate the weight of carbon dioxide obtained from a given weight of any one of these forms.

5. Prepare carbon dioxide by treating chalk or carbonate of soda (washing soda) with an acid. Prove that the gas thus obtained really contains carbon by heating a pellet of potassium in the dry gas contained in a small flask.

6. Demonstrate the high specific gravity of carbon dioxide by pouring it from one vessel to another, and showing that it extinguishes a taper.

7. Pass carbon dioxide over red-hot carbon in an iron tube, and show that it loses a part of its oxygen and is converted into carbon monoxide, a combustible gas, which, on combustion, again yields carbon dioxide. Collect the carbon monoxide over water containing caustic soda, and show that the gas does not render lime-water turbid. Then burn it, and show that the residual gas does possess this power.

8. Pass carbon monoxide over red-hot copper oxide to show the formation of carbon dioxide, and explain the use of carbon monoxide as a reducing agent in metallurgical operations.

9. Explain the changes which take place in an ordinary coal fire. Mention the poisonous nature of the carbon monoxide, and state that it is formed in cases of incomplete combustion from insufficient supply of oxygen.

10. Mention heat of combustion of carbon, and of carbon monoxide, and explain the value of the latter as a fuel.

11. Explain the reaction which takes place when carbon dioxide is passed into caustic soda and into lime-water, and explain the formation of a soluble carbonate in the first, and an insoluble carbonate in the second case.

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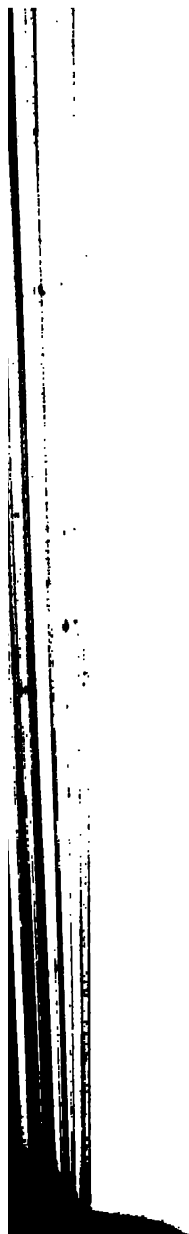
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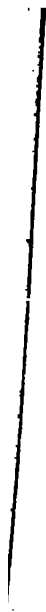
























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